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AN

INTRODUCTION

TO

SCIENTIFIC CHEMISTRY;

DESIGNED FOR

THE USE OF SCHOOLS

AND

CANDIDATES FOR UNIVERSITY MATRICULATION EXAMINATIONS.

BY

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PREFACE.

The method adopted in this book is somewhat different from that of other elementary treatises. The subject, which extends only to the *systematic* consideration of what are called the "non-metallic elements," indirectly treats of the more common metals which enter into combination with them, and is divided into two parts. In Part I. no symbols or atomic weights are used, but definite quantities in grammes are employed. I have always found that beginners have great difficulty in conceiving abstract numbers, and that symbols and equations are regarded by many as almost insuperable impediments; the attempt to overcome which involves a considerable amount of labour, which might be much more usefully applied.

It has not been my wish to make the subject what is called "easier," or to lessen the attention requisite for

acquiring it, but to direct that attention in a way which will render its exercise more profitable. Only a few theoretical considerations are introduced, and these have been reserved entirely for Part II., when the student is supposed to have acquired a knowledge of the facts on which those considerations are founded. Simple examples of calculation, and questions on the matter of each chapter, are given; and the student is earnestly requested to work out the one, and carefully write out the answers to the other, without referring to the book, and afterwards to compare them with the book, or submit them to his teacher. As only a few can acquire a sound knowledge of Chemistry by themselves, I by no means consider that this book should supply the place of the teacher. Some considerable experience in teaching young persons has proved to me that Chemistry cannot be learnt (it may be "crammed") without experimental illustrations and explanations given by a competent teacher; nor is even this sufficient without the student be himself exercised in practical work.

I am informed that Professor Bunsen, of Heidelburg, does not allow formulæ or atomic weights to be used during the first three months of a young student's course. It is gratifying to me to find that I have such confirmation of the course which I have for some time adopted.

A clergyman, who was a pupil of Gay Lussac, tells me that that famous chemist never used formulæ, or even mentioned the atomic theory, till he had fully treated of elementary bodies and their compounds; and that he gave as his reason for so doing, that facts are incontrovertible, but theories may be changed or modified; and therefore a knowledge of facts should be acquired before theories are considered.

Whilst intending this book to be a manual for candidates for the London University matriculation examination, I have not entirely confined myself to the curriculum at present put forth by that institution, as I hope that it will be found useful generally in schools, and in preparing students for other examinations, as well as in laying a sound foundation for the scientific study of Chemistry. With this view I have endeavoured, as much as possible, to make it an introduction to "Chemistry for Students" by Professor Williamson, whose assistant I have the honour to be, and to whom I am indebted, in great part, for what knowledge of Scientific Chemistry I possess, as well as for many personal favours. I am also indebted to Mr. Temple A. Orme, Teacher of Chemistry in University College School, and to Mr. J. J. Bowrey, late of University College, who have given me many valuable suggestions, which I have adopted; and to the Rev. Stephen Williams, Professor of Chemistry, Stonyhurst College,

and to Mr. J. Jones. Assistant to Professor Williamson, who have helped me with the calculations, and in revising the proofs.

FREDK. S. BARFF.

UNIVERSITY COLLEGE, LONDON,
March, 1869.

PREFACE TO THE SECOND EDITION.

IN this edition typographical and other errors have been corrected; no new matter has been added, except the chemical questions of the Matriculation Examination of the London University for June, 1869, together with the answers to those which involve calculation. *The nomenclature adopted in this book is the same as that now used by the Examiners at the London University.* In the last matriculation paper the terms *hydrogen* sulphate and nitrate are employed, instead of *hydric* sulphate and nitrate; this, however, is unimportant. The fact that these bodies are *salts* is recognised, and this being so, till some other more suitable word is used to designate them, SO_3 and its analogues must be called *acids.*

October, 1869.

INTRODUCTION.

The present state of chemical science, and the differences of opinion which exist on many important points in connection with it, render it difficult in an elementary work to use terms which will be considered unexceptional by all. Nor is it desirable to raise disputed questions where what is asserted will have, to a great extent, to be received by the student on trust, until his knowledge enables him to grapple with difficulties and form his own opinions. The nomenclature which has been employed was systematised by Professor Williamson and is used, to a considerable extent, by Professor Roscoe in his admirable little book, and appears likely to become general, founded, as it is, on a rational principle. The substantive use of the word "acid" is confined to those bodies which have for a short time been called anhydrides. What have been generally termed acids are admitted by all authorities to be hydrogen salts, and they are designated as such whenever it is found necessary to allude to them.

As the elements of heat are not treated of in this book, the conditions under which *all* volumes of gases and vapours are considered, are those of the standard temperature and pressure, *i.e.* at the temperature of zero Centigrade and under 760 millimètres of mercury pressure. The young student generally finds considerable difficulty in understanding how the vapour of water (steam) can be measured under such circumstances. It may be well to give as simple an illustration of it as possible. Gases when heated expand, and the rate of their expansion is regular, or very nearly so. If a measure of gas be heated to 1 degree Centigrade, its volume will increase $\frac{1}{273}$. Now this increase is at the same rate whether the measure be large or small. If then a measure of gas be heated from zero to 273°C. its volume will be doubled, thus—1 volume at zero Centigrade, when heated up to 1°C., becomes $+\frac{1}{273}$, when heated to 2°C. it becomes $1+\frac{2}{273}$, to 3°C. $1+\frac{3}{273}$, and so on to 273°C., when it becomes $1+\frac{273}{273}$, or $1+1$, *i.e.* 2 volumes; now the volume of 2(11.2) litres of steam at 100°C. is = 30.605 litres. So that, when it is said that 18 grammes of water in the state of steam occupy 2(11.2) litres, it does not mean that they occupy this volume at the moment of their formation by the union of oxygen and hydrogen, for at that time great heat is evolved and their volume is greater than 30.605 litres; but it

means that if the volume of the steam could be reduced to zero Centigrade, whilst the barometer indicated a pressure of 760 millimètres of mercury, then, under those circumstances, the steam would occupy 2(11.2) litres. Suppose at the time of the union of oxygen and hydrogen in the eudiometer, provision were made for heating it, so that the steam formed would be at 100°C., and if it were then found to measure 30.605 litres, its supposed volume at zero would be ascertained as follows : 273 volumes at zero become 373 volumes at 100°C., therefore, inversely, 373 : 273 : : 30.605 : 22.4. But, whatever be the volume of 18 grammes of steam, it is always the same as the volume of 2 grammes of hydrogen under the same circumstances; for it is manifest that if steam obeys the same law as fixed gases, under whatever temperature, and under whatever pressure, hydrogen unites with oxygen to form steam—the steam will occupy the same volume as the hydrogen did before its union with oxygen. This will be better understood by the student after he has read to the end of Chapter VII. and should be then referred to.

As it would be injudicious in an elementary work to introduce anything which has not received the public sanction of acknowledged authority, the Author has only stated, to the best of his ability, those opinions which he believes to be generally received. The lumi-

nosity of flame he has explained as owing to the presence in it of solid matter; he feels it, however, his duty to state that Dr. Frankland is at present engaged in investigating this subject, and from the able experiments, the results of which he has in part published, it is probable that "the solid particle theory" will have to be modified, if not abandoned. The atomic theory is briefly explained without any allusion to differences of opinion which exist respecting it.

The specific gravities of liquids and solids, and of gases referred to atmospheric air, are only occasionally alluded to in the text, but a table of those of the most important elements is given at the end of the book.

The decimal system of weights and measures is employed, and a table comparing them with English weights and measures is appended.

The difference between the Centigrade and Fahrenheit thermometric scales is given in Chapter I.; but afterwards all temperatures are recorded in degrees Centigrade.

By showing young persons the benefits to be derived from the use of a rational metrical system, it is to be hoped that they will, in their time, assist in getting it generally adopted in this country.

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PART I.

PRELIMINARY CONSIDERATIONS.

CHAPTER I.

Chemistry treats of the changes which take place when one substance acts upon another so as to form new bodies, having different properties from either of the substances employed. In order that substances may act chemically on one another, the particles of which they are composed must come into, what is ordinarily called, actual contact. Sometimes two solid bodies may be mixed together in a mortar, as, for instance, sodic carbonate, commonly called soda, or carbonate of soda, and hydric tartrate, commonly called tartaric acid, and if both be dry no action takes place; but if water be added, effervescence occurs, bubbles are formed, and a gas escapes, and a new substance is left behind, which is altogether different from either hydric tartrate or sodic carbonate. If these substances have been mixed in proper proportions, the gas which escapes is carbonic acid, and the substance which remains dissolved in the water is sodic tartrate. Here the water has brought into close contact the particles of the two substances. Again, if fine copper filings be mixed with finely-powdered brimstone—the chemical name is sulphur—there will be a mixture of copper and sulphur, small

pieces of red copper with small pieces of yellow sulphur, and the sulphur and copper can be completely separated from one another: but if they be heated together, a new substance is formed, of a black colour; no copper remains as such, if the heat has been applied long enough; the heat has melted the sulphur and has brought the particles of sulphur and copper into close contact, and chemical action has taken place, and cupric sulphide has been made. You will be expected to know the difference between a mechanical mixture and a chemical compound, and this is a good illustration of it; when the copper and sulphur were first mixed, and until heat was applied, the mixture was called a mechanical mixture; but after the action of heat it became a chemical compound. In like manner the sodic carbonate and hydric tartrate, as long as they were kept dry, formed a mechanical mixture, but when water had been added, chemical action took place, and new chemical compounds were formed. We will take one more illustration. If sodic carbonate be mixed with sand, no change takes place, and the sodic carbonate can be washed away entirely from the sand; but if they be heated together, carbonic acid escapes and a kind of glass is formed. This is a chemical compound quite different from sand or sodic carbonate.

Simple bodies which are believed to consist of only one substance are called elements, such as iron, silver, gold, charcoal, sulphur, &c.

Compound bodies are made up of two or more elements chemically combined, as cupric sulphide.

Some substances are more dense than others; that is, their particles are heavier. For instance, take two

marbles of exactly the same size, one of clay, the other of stone or marble; that of stone or marble is the heavier of the two, that is, it is more dense. If quicksilver be poured into water it sinks to the bottom, because it is more dense than water; and oil floats on the surface of water, because it is less dense than water.

Again, some gases are more dense than others Common air is a mixture of two gases. And the gas (coal gas) which we burn is also a mixture of gases; but this gas is lighter than air, as is proved by the fact that a balloon filled with coal gas will rise in the air. If you take two thin india-rubber air-balls of equal size, and fill one with air and the other with coal gas, that filled with air will fall to the floor, while that filled with coal gas will go up to the ceiling. In equal measures of coal gas and air, the coal gas is lighter than the air. There are other gases besides those we have mentioned which are heavier than air. Now it is necessary that we should be able to compare the densities of different bodies, and in order to do this we take one body as a standard, and then find out how much heavier other bodies are than it is: we take water as the standard for solids and liquids.

Thus, if we weigh a measure of water and then weigh a measure of iron of exactly the same size, we shall find that the iron is nearly eight times as heavy as the water. We shall not, however, consider the relative weights of solids and liquids here, as it belongs to what is usually considered another branch of the subject; but we shall have to consider the densities of gases and vapours, and to do so properly, we must at the commencement settle upon a definite standard of

weight and measure. We shall take the lightest gas known, viz. hydrogen, as our standard of reference, and with it we shall compare all other gases and vapours. The quantity by weight of this gas which we shall take as our standard will be one gramme, and the quantity by measure will be eleven decimal two litres—and the reason why this weight and this measure are taken is, because that quantity which will fill a measure containing 11.2 litres, weighs 1 gramme at the temperature of melting ice, that is zero on the Centigrade scale, and under the pressure of a column of mercury 760 millimètres high, and this is called the standard temperature and pressure.

The gases and vapours must always be weighed at the same temperature, because heat causes gases to expand, as may be clearly seen if a bladder about two-thirds full of air be warmed before a fire; before warming, the bladder, not being full of air, is loose, but when it is warmed the air expands and fills it, and the bladder becomes quite tight. The air has been made by the heat to fill a larger space, but its weight is the same as it was before heating. If therefore we were to let off half the air in the bladder, when heated, into a vessel and measure it, keeping it all the time at the same temperature, and if we were to allow it to cool down, we should find that it would, when cool, occupy a smaller space; and it would therefore weigh less than it did at the original temperature.

When gases are compressed they occupy a smaller space than they did before; thus, in the air-gun, air is pumped into a metal ball, and prevented from escaping by a stopcock—when the stopcock is turned the air

rushes out violently, and the air pumped into the small metal ball would, if allowed to escape into a vessel where it could be measured, measure very much more than it did when in the ball.

Now the atmosphere which surrounds the earth extends to about 44 miles above its surface, and, consequently, the air near the earth, being pressed upon by the air above it, is more dense than the air higher up, which has less air above it—so that a pint of air on the surface of the earth weighs heavier than a pint of air on the top of a high mountain. Suppose a column of air 1 inch square and 44 miles high, that is as high as the air extends from the earth in an upward direction, could be separated out from the atmosphere and weighed, it would weigh about 15 lbs. Therefore it presses on a piece of the earth's surface of 1 inch square with a pressure of 15 lbs. If we take a glass tube about a yard long and with an area equal to 1 square inch, and seal up one of its ends, and, after filling it with mercury, place it with its open end in a vessel containing mercury, the mercury in the tube will fall till the top of the column in the tube is 30 inches above the surface of the mercury in the vessel—and this is because a column of air 44 miles high and 1 inch square is pressing upon the lower end of the column of mercury, and as long as the column of mercury weighs as much as the column of air, the air can support it in the tube. Now this column of mercury 30 inches high is found to weigh 15 lbs.—and therefore we know that the pressure of the column of air is 15 lbs. If any circumstances cause the column of air to be lighter than 15 lbs., then the column of

mercury will be less than 30 inches; and if the column of air is brought to weigh more than 15 lbs. then the column of mercury in the tube will be longer than 30 inches. Now such circumstances do occur, and therefore the air is more dense at one time than at another. (This we know from the barometer; sometimes the mercury in it stands high, and then the air is dense; sometimes it stands low, and this shows that the air is light.) And therefore gases which are pressed upon by the air will be more dense when the air which presses on them is heavier or more dense; and if a gas is more dense at one time than another, the same measure of it will weigh heavier. We see, therefore, that it is necessary to measure gases under the same pressure.

A simple method of weighing gases is to get two glass globes of equal capacity, and of as nearly equal external dimensions as possible. These globes should be supplied with stopcocks. Both should be exhausted of air by the air-pump, and should be made to weigh exactly the same weight in a very delicate balance. One of these globes should be removed from the balance and be filled with hydrogen and weighed, and its weight noted carefully. It should then be emptied, exhausted, and filled with, say, oxygen, and weighed again; this time it will be found to weigh 16 times as much as it did before, and in this way the relation between the density of any gas and hydrogen can be obtained, or that existing between the densities of any two gases.

The temperature at which gases are measured is zero Centigrade. There are three kinds of thermo-

meters by which the temperature is measured,—Fahrenheit, which is used in this country, Centigrade, and Reaumur, which are used on the continent. The Centigrade thermometer is almost invariably employed in scientific experiments, and is generally quoted in scientific works; we shall therefore use it in this book. All thermometers are marked with divisions which are called degrees. In a Fahrenheit thermometer the point called freezing-point is at 32°, and a point 32° below this is called zero. The freezing-point in a Centigrade thermometer is called zero; the zero of Centigrade, therefore, corresponds with 32° of Fahrenheit. In a Fahrenheit thermometer the boiling-point is 212°, but in the Centigrade it is 100°. Now, between the freezing-point in Fahrenheit and the boiling-point there are 180 degrees, because there are 212 degrees from the zero, and this is 32 degrees below the freezing-point, and 32 subtracted from 212 give 180. In Centigrade the number of degrees between the freezing and boiling points is 100, therefore, in the two thermometers the same space is divided into 180 and 100 parts respectively, consequently a degree Centigrade is larger than a degree Fahrenheit, in the proportion of 180 to 100, or 9 to 5. We can find out to what degree Centigrade any given degree of Fahrenheit corresponds if we first subtract from it 32, because the zero of Fahrenheit is 32 degrees below the freezing-point, and then multiply the remainder by 5 and divide by 9.

$$C = (F - 32)\tfrac{5}{9}$$

And if we wish to change degrees Centigrade into degrees Fahrenheit we must multiply by 9 and divide by 5, and then add 32.

$$F = C\tfrac{9}{5} + 32.$$

For example, to what degree Centigrade do 212° Fahrenheit correspond?

$$C = \frac{(212 - 32)\,5}{9}$$

$$= \frac{180 \times 5}{9}$$

$$= \frac{900}{9} = 100°$$

Or, to what degree Fahrenheit do 100° C. correspond?

$$F = \frac{100 \times 9}{5} + 32$$

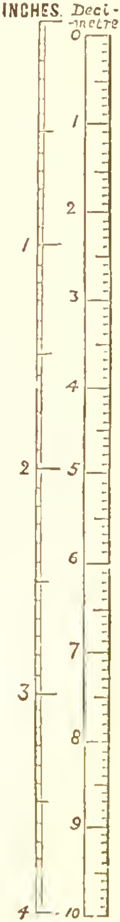
$$= \frac{900}{5} + 32 = 180 + 32 = 212°$$

All degrees below zero Centigrade are marked with the minus sign (—) before them thus, —5° means five degrees below freezing-point.

The pressure of the atmosphere at which gases are measured is that of 760 millimètres of mercury, and this is equal to a pressure of 15 lbs. on the square inch, or of a column of mercury 30 inches high.

The weights and measures which we shall employ are those of the decimal system; they are used in all foreign scientific works, and very generally in all new books on such subjects published in this country. They are also much more convenient and rational than our own. In this system the standard measure of length is called a

mètre; it occupies the place of our yard, but is nearly $3\frac{1}{3}$ inches longer than the English yard. The mètre is the 10,000,000th part of the length of a line drawn directly from the earth's equator to the pole. A standard mètre is kept in Paris for reference; its exact length is 39.37079 inches. The mètre is divided into ten parts called decimètres, and each of these again into ten other parts called centimètres, because each of these second divisions is 100th of a mètre, and these centimètres are divided into millimètres because each of them is 1,000th of a mètre. The accompanying figure represents a decimètre divided into centimètres and millimètres. A kilomètre consists of 1,000 mètres, and is about two-thirds of a mile. The decimètre is 3·937 inches. A cube whose sides measure each 3·937 inches is a cubic decimètre, and this is taken as the standard measure of capacity, just as the pint is the English standard; but the cubic decimètre contains more than a pint; it holds 1·765 pints, and it is also called a *litre*. A litre therefore, in round numbers, holds about $1\frac{3}{4}$ pints, or a little more. The litre is divided into decilitres, *i.e.* tenths of a litre, and into centilitres, *i.e.* hundredths of a litre. The standard of weight is the weight of a cubic centimètre of distilled water at the temperature of 4° C. This weight is called a gramme, and is equal to 15·432349



grains English weight. The gramme is divided into decigrammes or tenths, eentigrammes or hundredths, and milligrammes or thousandths of a gramme. A weight of 1,000 grammes is ealled a kilogramme, and it weighs 2·2046213 pounds avoirdupois, or rather less than $2\frac{1}{4}$lbs.

CHAPTER II.

HYDROGEN.

Hydrogen is a colourless gas, that is, it is clear and transparent like air, it has no smell, and it is the lightest body known. It enters largely into the composition of water, and is also found in small quantities uncombined; it exists in coal, and from it is obtained, united with another substance forming coal gas, the gas we burn. It is also a constituent of most animal and vegetable bodies.

Hydrogen is obtained from some substance which contains it. The substance usually taken for its production is one which consists of three elements—sulphur, hydrogen and oxygen. This substance is a liquid, and when mixed with water and zinc, effervescence takes place, and hydrogen escapes. This gas can be collected in vessels. The sulphur and oxygen together, unite with the zinc and form a new substance in which zinc takes the place of the hydrogen which has been given off. The substance which is mixed with the water and zinc is hydric sulphate, and is commonly called sulphuric acid, and its commercial name is oil of vitriol. When zinc acts upon it, the zinc drives off the hydrogen because the metal zinc is

chemically more powerful than the metal hydrogen (for hydrogen is a metal, existing in nature as a gas; this you will understand better later on); and as hydric sulphate was originally taken, and as zinc has taken the place of hydrogen, which has gone away, zinc sulphate must remain.

Water is added to the zinc and hydric sulphate, because zinc sulphate which is formed will not dissolve in hydric sulphate, but it dissolves readily in water,—and it is necessary that the zinc sulphate should be dissolved, or it would prevent the action continuing. 98 grammes of hydric sulphate contain 2 grammes of hydrogen, 32 grammes of sulphur and 64 grammes of oxygen, and if this be acted upon by 65 grammes of zinc, the 2 grammes of hydrogen are set free, and the 32 grammes of sulphur and 64 of oxygen unite with the 65 grammes of zinc to form 161 grammes of zinc sulphate.

As the vessel in which the hydrogen is generated contains water, the hydrogen will take up moisture, and hold it in suspension. This is the case with all gases which have been in contact with water; in order to dry them they must be passed through a tube containing a substance called calcic chloride, which has been melted and subjected to a red heat so as to make it perfectly dry. Or through a bottle containing hydric sulphate (lumps of pumice-stone are usually placed in this bottle, soaked with the sulphate so as to expose a larger drying surface to the gas); both these substances take up moisture very readily.

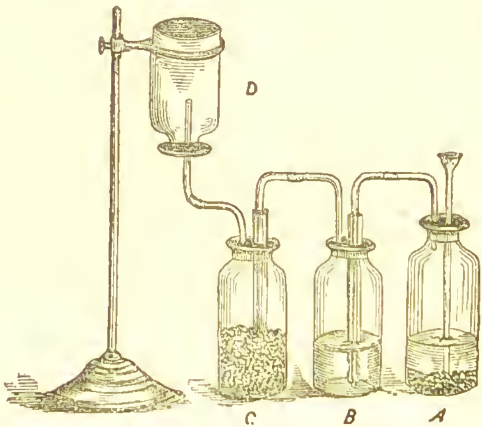
As the hydric sulphate used in the preparation of hydrogen almost always contains impurities they will

affect the purity of the gas, because they form chemical compounds with it, which are gases, and which will pass over in its company. These impurities, which are arsenic, sulphur and perhaps phosphorus, can be got rid of by passing the gas through a solution of potassic or sodic hydrate, or of argentic nitrate, or of both taken in succession,—this purifying process must evidently take place before drying the gas with hydric sulphate or calcic chloride. Iron can be used instead of zinc for the decomposition of the hydric sulphate, but the hydrogen produced in this way has an unpleasant smell owing to the presence of carbon, as an impurity, carbon existing in ordinary cast-iron.

Other bodies containing hydrogen can be used for its preparation; for example, potassic hydrate, which contains potassium, oxygen and hydrogen, when acted on by the same metal, zinc, loses its hydrogen which is set free, zinc taking its place, so that potassic hydrate becomes potassic zincate. Water also contains oxygen and hydrogen; when water is boiled steam is formed, and if this steam be passed through a tube containing very small particles of iron heated to redness, the iron, like the zinc in the former cases, turns out the hydrogen from the water, in the state of steam, and the hydrogen is set free, the oxygen of the water uniting with the iron. As water, the compound of hydrogen and oxygen, is called by chemists hydric oxide; when the iron takes the place of the hydrogen, the substance becomes a *ferric* oxide (ferric is derived from the Latin word *ferrum*, iron). This action will be again referred to when the exact composition of this oxide will be given. If the metal

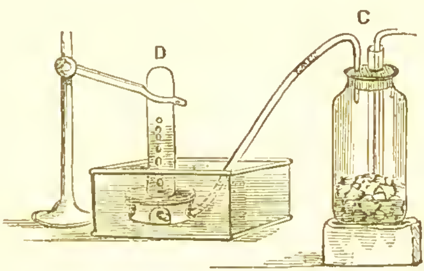
sodium be thrown into water, a violent action takes place,—hydrogen is set free, and the sodium unites with the oxygen, and part of the hydrogen of the water and sodic hydrate remains,—it is called sodic hydrate, because it contains sodium, oxygen and hydrogen. If instead of sodium, potassium be employed, the action is more violent, and the heat produced by the union of potassium with oxygen and hydrogen to form potassic hydrate is so great, that the hydrogen, which is set free, is ignited, and burns with a beautiful violet coloured flame.

You see then, that when we want to get free hydrogen by itself, we must take some compound which contains it, and we must turn out the hydrogen by some metal which, under the circumstances at the time, is more powerful than it is. Hydrogen, we have said, is the lightest body in nature. If we fill a vessel containing 11.2 litres with hydrogen and weigh it, it will be found to contain 1 gramme of hydrogen. 11.2 litres of air weigh 14.45 grammes, it is therefore about 14 times heavier than hydrogen.



When we prepare hydrogen, we usually take hydric sulphate, zinc and water (because this is the most convenient method), and we put them into a glass bottle *A*, with a tube fitted into its cork to allow the escape of the gas. As the gas escapes, it, being lighter than air, ascends, so that if we hold (*D*) a glass vessel with its mouth downwards, over the escape pipe, the hydrogen will go up into the vessel and drive out the air, and the vessel will soon be filled with hydrogen —this is called collecting the gas by upward displacement. The vessels *B* and *C* in the diagram are for purifying and drying the gas. *B* contains a solution of sodic hydrate or of argentic nitrate to absorb the impurities—arsenic, sulphur, &c., and the vessel *C* is in part filled with lumps of pumice-stone moistened with hydric sulphate to absorb water vapour.

Hydrogen can also be collected over water, that is, the end of the escape tube being put under water, bubbles will rise to the surface of the water; these are bubbles of hydrogen.



Now if a bottle (D) full of water be placed with its mouth in the water, where the escape tube is inserted and exactly over the end of it, the bubbles of hydrogen will rise through the water in the bottle,

and gradually drive it out, so that in a short time the bottle will be filled with hydrogen. Many, but not all, gases can be collected in this way, so that in future, when we speak of a gas being collected over water, you will remember that we allude to the method just described.

Hydrogen burns with a pale blue flame, which gives little or no light; although the hydrogen flame gives little or no light it is intensely hot.* Some gases, when collected in a vessel which is made extremely cold, become liquid, and others can be made liquid by great pressure, or by cold and pressure. The method of making gases liquid will be described later on,—but *hydrogen* you must remember has *never* been made liquid.

QUESTIONS—HYDROGEN.

1. How is hydrogen usually prepared? What are its impurities? How can they be removed?

2. What bodies in nature contain hydrogen?

3. Give the physical properties of hydrogen.

4. How many grammes of hydrogen are contained in 44.8 litres at the temperature of 0°C. and 760 mm. pressure?—*Ans.* 4 grammes.

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* Before applying a light to hydrogen gas, it is necessary to ascertain that it has no atmospheric air mixed with it (for if hydrogen mixed with air be lighted, a violent explosion takes place); this is done by collecting some of the hydrogen in a small test tube, and lighting it; if it lights quietly the gas is free from air, but if it makes a sharp shrill noise on application of the light it is not safe to light the gas coming from the generating bottle.

5. What is the action of zinc on potassic hydrate?

6. Describe a method of replacing hydrogen in water by iron?

7. What weight of hydrogen can be evolved from 294 grammes of hydric sulphate?—*Ans.* 6 grammes.

8. How much hydric sulphate is required to evolve 145.6 grammes of hydrogen?—*Ans.* 7134.4 grammes.

9. A balloon is capable of containing 5,600 litres of gas, what quantity of hydric sulphate must be acted upon by zinc to evolve sufficient hydrogen to fill it?—*Ans.* 24,500 grammes.

10. When hydrogen is prepared from zinc and hydric sulphate what substance remains after the hydrogen has all escaped? When steam is passed over red-hot iron, what compound is produced with the iron?

11. When hydric sulphate is acted upon by 130 grammes of zinc, what weight of hydrogen is replaced by the zinc?—*Ans.* 4 grammes. How many litres would the gas occupy at the standard temperature and pressure?—*Ans.* 44.8 litres.

12. State fully the action of potassium on water, and in what does this action differ from that of sodium?

CHAPTER III.

OXYGEN.

Oxygen is a colourless transparent gas, without smell; it exists free in the atmosphere; it forms about one-fifth of its bulk; that is, in a vessel containing 5 litres of air there are 1 litre of oxygen, or thereabouts, and 4 litres of a gas called nitrogen. Oxygen also exists combined with metals, as, for example, in ferric oxide, common iron-rust; it also forms part of many compounds, both animal and vegetable, as wood, sugar, the flesh and fat of animals, &c., and is so widely distributed in nature that it is said to constitute one-half the weight of the earth's crust. It also forms the greatest part, by weight, of water.

Oxygen is usually obtained from substances which contain it chemically combined; though it is also got from the air where it is free, that is, where it does not exist in chemical combination. Oxygen was discovered at about the same time—about 1775—by Dr. Priestley, in England, by Scheele, and by Lavoisier, in France. It was first obtained from the atmosphere by heating quicksilver (its chemical name is mercury) in air. The mercury was observed to change its appearance when heated in air, and the air in the vessel in which the mercury was heated was found to lessen in bulk, and the mercury to gain in weight. When the vessel with the air in it was weighed, the weight which it had lost was found to correspond with the weight which the mercury had gained; and when afterwards the mercury was heated

again to a higher temperature than before, and the products carefully collected, the substance given off from the mercury was found to be a gas, and to measure just as much as the air in which the mercury had been heated had lost in bulk. When this gas was returned again to this air, it was found to restore to it certain properties which the air had lost after the mercury had been heated in it.

Potassic chlorate is the substance which is generally used for obtaining oxygen. Potassic chlorate contains potassium, chlorine and oxygen. When potassic chlorate is heated, it melts, and then seems to boil; bubbles arise, and these are bubbles of oxygen. If the potassic chlorate be carefully and sufficiently heated, all the oxygen which it contains will be given off. 122.5 grammes of chlorate contain 39 grammes of potassium, 35.5 grammes of chlorine, and 48 grammes of oxygen; so that from 122.5 grammes of potassic chlorate we can get 48 grammes of oxygen. Now, suppose we want to find out how much oxygen we can get from 50 grammes of potassic chlorate, we have to make a simple Rule of Three statement:—

Grms. of chlorate.		Grms. of oxygen.		Grms. of chlorate.
122.5	:	48	: :	50

```
                 48
                 50
              -------
 122.5 )      2400   ( 19.59
              1225
              -----
              11750
              11025
              -----
              ..7250
                6125
                -----
                11250
                11025
                -----
                ..225
```

Or, if you prefer it, you may state it in the following manner :—

Grms. of chlorate.		Grms. of chlorate.		Grms. of oxygen.		Required quantity of oxygen.
122.5	:	50	::	48	:	x

$$\frac{122.5}{50} = \frac{48}{x}$$

$$x = \frac{50 \times 48}{122.5} = 19.59$$

But suppose you wish to know how much potassic chlorate you must take to make 50 grammes of oxygen, the statement will be, either—

Grms. of oxygen.		Chlorate.		Grms. of oxygen.
48	:	122.5	::	50

	Oxygen.		Oxygen.		Chlorate.		Required chlorate
or,	48	:	50	::	122.5	:	x

$$\frac{48}{50} = \frac{122.5}{x}$$

$$x = \frac{122.5 \times 50}{48} = 127.604 \text{ grms.}$$

After the potassic chlorate has lost all its oxygen, 39 grammes of potassium and 35.5 grammes of chlorine will remain chemically combined, forming 74.5 grammes of a substance called potassic chloride: and if it is desired to know how much potassic chloride is

obtained from 50 grammes of potassic chlorate, you have simply to work out the following statement:—

Grms. chlorate.		Grms. chloride.		Grms. chlorate.		
122.5	:	74.5	::	50		
or, 122.5	:	50	::	74.5	:	x

All calculations for determining the quantity of any constituent of a chemical compound are done in the same manner as these, and are all equally simple. In the examples above, we have taken our weight of potassic chlorate in grammes because we have arranged to work in grammes for the present; and we have taken 122.5 grammes of that substance for reasons which you will very soon begin to understand. And we show by *experiment*, that 122.5 grammes of potassic chlorate contain 39 grammes of potassium, 35.5 of chlorine, and 48 of oxygen.

Oxygen is also prepared from manganic binoxide, called also black oxide of manganese. When this substance is heated to a very high temperature, it gives up only one-third of its oxygen, a lower oxide of manganese being left, so that only *one-third* of the oxygen which it contains can be got by heating manganic binoxide.

Oxygen can also be obtained from other substances; these will be noticed in the second part of this book.

Oxygen is heavier than hydrogen. If the 11.2 litre measure be filled with oxygen and weighed, the oxygen contained in it, *i.e.* 11.2 litres of oxygen, will be found to weigh 16 grammes; and as 11.2 litres of hydrogen weigh 1 gramme, oxygen is 16 times as heavy as hydrogen; that is, if the density of hydrogen is 1, that

of oxygen is 16. As we know the volume or measure which 16 grammes of oxygen occupy, we can easily determine the volume or weight of the gas; say, for example, 50 grammes:—

Grms. of oxygen.		Litres.		Grms. of oxygen.		
16	:	11.2	::	50		
or, 16	:	50	::	11.2	:	x

$$x = \frac{50 \times 11.2}{16} = 35 \text{ litres.}$$

Or we can find what weight of oxygen is required to fill any desired measure; say 50 litres.

Litres.		Grms. of oxygen.		Litres.		
11.2	:	16	::	50		
or, 11.2	:	50	::	16	:	x

$$x = \frac{50 \times 16}{11.2} = 71.428 \text{ grms.}$$

And by combining the two calculations, viz. that of the quantity of potassic chlorate requisite to yield a desired weight of oxygen, and that for the discovery of the quantity of oxygen required to fill a given measure, we can find out how much potassic chlorate must be taken to give 100 litres of oxygen, the statement—

Litres.		Grms. of oxygen.		Given litres.
11.2	:	16	::	100

will give the weight of 100 litres of oxygen; and, calling this weight y, the following statement will give the quantity of potassic chlorate required to yield it—

Grms. of oxygen.		Grms. of chlorate.		Given grms. of oxygen.
48	:	122.5	::	y

When hydrogen is lighted in air it burns; but we cannot light oxygen in air at all—it will not burn in

atmospheric air. We have before said that oxygen is free in the air, that is, it is not chemically combined with any other substance; now, it is because the oxygen in the air is free, that hydrogen can burn in air, and the reason why hydrogen burns is because there is free oxygen for it to unite with. (There are circumstances in which hydrogen can burn in oxygen which is not free; these will be explained afterwards.) The oxygen then causes the hydrogen to burn, and it is therefore said to support combustion.

If a wood splint, which continues to glow after the flame has been extinguished, be plunged into a vessel of oxygen, the flame is immediately re-kindled. Phosphorus, iron, sulphur, and many other bodies burn very brilliantly in oxygen. A candle and a fire burn in air by virtue of the oxygen which the air contains. If a candle be burnt in a closed glass vessel, it will continue alight for a short time, and will then go out because all the oxygen has been used up. Oxygen diluted with four times its bulk of nitrogen also supports animal life; without it we should die: there is no other gas, elementary or compound, in which we could live. If a small animal be placed in a closed bottle, it will live as long as a certain quantity of oxygen remains free; but when this oxygen has become combined with other substances, so as to form other gases, the animal dies.

QUESTIONS—OXYGEN.

1. If mercury be heated in air, what compound results? If this compound be heated to a higher temperature, what gas is given off?

2. How is pure oxygen best obtained? and how many grammes of it occupy 67.2 litres at the standard pressure and temperature?—*Ans.* 96 grammes.

3. A vessel when empty weighs 100 grammes, when filled with hydrogen it weighs 102 grammes, and when filled with oxygen 132 grammes, what is its capacity?—*Ans.* 22.4 litres.

4. What weight of oxygen can be expelled from 100 grammes of potassic chlorate?—*Ans.* 39.183 grammes.

5. A gasometer has a capacity of 112 litres, how much potassic chlorate must be taken to fill it with oxygen at the standard pressure and temperature?—*Ans.* 408.33 grammes.

6. What part does oxygen play in the economy of nature?

7. How much oxygen can be obtained from a weight of manganic binoxide which contains 30 grammes of oxygen?—*Ans.* 10 grammes.

CHAPTER IV.

NITROGEN.

Nitrogen is a colourless transparent gas. It has no smell; it does not support life, *i.e.* animals and vegetables die when kept in an atmosphere of it; a candle will not burn in it, and it is altogether the most inactive of gases. It forms four-fifths of the air, the other fifth being, as already stated, oxygen. It exists largely in nature combined with other elements in animal and vegetable substances. It is also found in a substance called nitre, or saltpetre, and from this its name nitrogen is derived.

Nitrogen can be obtained from atmospheric air. If the oxygen in the air be taken away from it, nitrogen will remain. A very convenient way to get nitrogen is to pass pure air over red-hot very finely-divided metallic copper contained in a glass tube; the copper takes the oxygen from the air, forming a substance called cupric oxide (cupric is derived from the Latin word *cuprum*, copper), and the nitrogen passes on and can be collected in a vessel. When steam was passed over red-hot iron, you will remember that the iron took the oxygen from the steam, forming a ferric oxide, and hydrogen passed over; here the copper

takes the oxygen from the air, and nitrogen passes over.

If phosphorus be burnt in pure air, its oxygen unites with the phosphorus, forming phosphoric acid, and *only* nitrogen remains as a gas. Or if a piece of phosphorus be suspended in a bottle which has its mouth placed in water, the phosphorus will slowly unite with the oxygen of the air contained in the bottle, and the phosphoric acid formed will be dissolved in the water, and the space which the oxygen formerly occupied in the bottle will be taken by the water, which will rise to fill one-fifth of the bottle, and the remaining space above the water will be filled with nitrogen gas.

Nitrogen can be obtained in other ways than those already mentioned; they will be given in later chapters, as a greater amount of chemical knowledge is required to understand the reactions than you are at present supposed to possess.

Nitrogen is heavier than hydrogen, but not so heavy as oxygen. If our 11.2 litre measure be filled with nitrogen, the weight of the nitrogen will be 14 grammes; therefore nitrogen is 14 times as heavy as hydrogen, that is, hydrogen being taken as 1, nitrogen is 14, and therefore the density of nitrogen is said to be 14.

Questions—Nitrogen.

1. How may nitrogen be prepared from air? What is its density?

2. How does it exist in nature?

3. If phosphorus be burnt in air enclosed in a vessel, what gaseous body remains?

4. What volume do 168 grammes of nitrogen occupy? —*Ans.* 134.4 litres.

CHAPTER V.

CARBON.

The three elements which we have been considering are all gases, carbon is a solid, the commonest form in which we find carbon is as charcoal—black charcoal. It is also very common in the form of black-lead: (there is no lead in what is called black-lead) of which drawing pencils are made, and which is used for blackening grates. Other names for this form of carbon are plumbago and graphite. Graphite is found in mines in Cumberland, also in Siberia and Ceylon; it occurs crystallised. It is a heavy powder, it looks like a metal, it can be made into hard masses, by being pressed very heavily—the commonest kinds of pencils are filled with this compressed powder, the best with strips cut out from the native lumps. Carbon also occurs in nature as diamond; the brilliant diamond is chemically the same as charcoal or graphite—but it is very pure carbon crystallised. It is very hard, the hardest substance in existence, and is used to cut glass, also for jewelry.

Carbon then exists in three states, which are called "*allotropic modifications.*" It exists in combination with other elements as gases, in all animal and vegetable substances. When wood or flesh is strongly heated

in a closed vessel a black substance remains; this is carbon—the black residue, which remains when any organic substance is heated sufficiently to decompose it, is almost entirely carbon. The colour called lamp-black, which is the smoke from some resinous substance, as turpentine, is carbon; coals, coke and soot also contain carbon.

Charcoal, which is wood heated out of contact with air, appears to be very light, but this is because it is very porous, *i.e.* full of holes; it has the peculiar property of absorbing gases, which are condensed in its pores, and for this reason if meat which smells strongly, owing to bad gases being given off from it, be placed in charcoal, it will soon become sweet. If bones of animals are exposed to a great heat in a closed vessel, the black residue is animal charcoal; mixed with some other substances, this is largely used for purifying bad water. If dirty water, which smells strongly, be passed through animal charcoal, it loses its bad smell and becomes quite clear and sweet.

The coal we burn contains large quantities of carbon together with hydrogen. When coal is burnt out of contact with air, as in a gas retort, nearly pure carbon remains; it is called coke, and is a sort of artificial graphite. The hydrogen in union with some carbon contained in the original coal goes off as gas, and is collected in gasometers for use.

Coal is formed from wood. Ancient forests which once grew on the surface of the earth became buried, and changes took place in the wood: a substance called bitumen, which yields tar, was formed; the structure of the wood was changed, and coal was

formed. Some coals contain more of this bitumen than others, and therefore they contain more hydrogen, and the hydrogen in coal exists in the bituminous matter. Cannel coal is of this kind: it yields more gas when heated, and blazes cheerfully in the fire-grate; while anthracite coal, which is nearly all pure carbon, does not burn brightly, but gives out a great amount of heat. Besides carbon, oxygen, and hydrogen, coal contains some sulphur and nitrogen, so that in gas-making compounds of hydrogen and sulphur, of carbon and sulphur, and of nitrogen and hydrogen are formed; these compounds will be studied in their proper place, when a more detailed account of the manufacture of coal gas will be given. Carbon has been subjected to the most intense heat which we have at our disposal, but it has never been converted into vapour; we cannot, therefore, ascertain *directly* its density with respect to hydrogen.

Note.—Questions on carbon will be given after the chapter on compounds of oxygen and carbon.

CHAPTER VI.

COMBUSTION.

We have seen that hydrogen burns in air, that is, in the oxygen which is contained in the air, and as this oxygen is diluted with nitrogen, the temperature produced is not so great as when hydrogen is burnt in pure oxygen. When hydrogen is mixed with oxygen in a vessel, the two gases remain separate; they do not unite, but if they be heated to a high temperature, they do unite and form a compound gas quite different from either oxygen or hydrogen. Their union is attended with increased heat, and so great is that heat, that it will melt a metal called platinum, which cannot be melted in the hottest furnace. Moreover, when these gases are confined in a vessel, their union is attended with a loud report.

When hydric sulphate (oil of vitriol) is poured into water, the vessel containing the mixture becomes very hot, the hydric sulphate unites chemically with the water, and this union causes heat to be made manifest to our senses. If hydric sulphate be poured upon common sugar, which is a compound of carbon, oxygen, and hydrogen, the oxygen and hydrogen

unite to form water, which again unites with the hydric sulphate and their chemical combinations cause heat to be liberated.

Many other examples of this evolution of heat attending chemical combination could be given, but it is better at present not to introduce substances with which we have no acquaintance, these given being quite sufficient for our purpose. In the first instance oxygen and hydrogen unite to form a new body and heat is evolved; in the second, hydric sulphate and water unite, forming a body different from either water or hydric sulphate; and in the last, hydric sulphate takes into union with itself the elements of water, as water, from the sugar, a chemical compound of hydrogen and oxygen, viz. water, being formed which did not exist in the sugar, and then a chemical compound of that water with hydric sulphate, rendering it chemically different from what it was before, and in both these combinations heat is manifested. Union of element with element, or of compound with compound, or of elements with compounds, is in every case attended with production of heat, and is what is commonly called *combustion.* When we speak of a body burning, we simply mean that some or all of its constituents are entering into chemical combination with some other bodies, to form new bodies different in their properties from those engaged in forming them. In some cases, burning or chemical combination cannot take place without the substances which are so combined being raised to a high temperature; our first example of the union of oxygen and hydrogen is an instance of this; but in other cases heat and

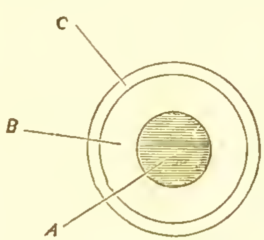
even light are manifested when substances are mixed together at the ordinary temperature of the air.

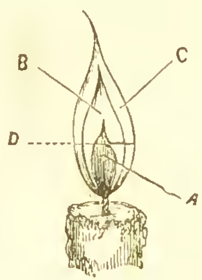
It has been said that animal bodies contain carbon, they contain it in several different combinations, and it is contained in the blood—in the dark blood which flows from veins. You know the difference between the colour of the blood which comes from a cut finger, and from a vein when a person is bled in the arm: one is bright red, the other, which flows from the vein, is of a dark brownish red colour. Now this dark blood passes through the lungs, which are in the chest, and the air which we inhale contains oxygen, and this oxygen meets the dark blood in the lungs, and it unites with the carbon in the dark blood, and forms a chemical compound of oxygen and carbon; and so a burning takes place in our bodies, when the oxygen and carbon unite and our bodies are warmed, and this is the principal source of animal heat.

If potassic chlorate and sugar be separately pounded and mixed together, and if a small quantity of hydric sulphate be poured on the mixture, it will immediately catch fire, and an explosion will take place similar to that which is produced when gunpowder is fired. The oxygen of the potassic chlorate is set free, and unites with the carbon and hydrogen of the sugar; the oxygen in the sugar also takes part in these reactions, and so intense heat and light are produced, which are the results of chemical combination. Again, if metallic iron in a very finely-divided state, which has been kept in a closed vessel, free from oxygen, be shaken out into the air, it will become quite red-hot and will take fire; and if the powder be caught on a sheet of white paper,

it will be found to be red, the original colour of the metal being nearly black, the powder is the red oxide of iron (ferric oxide); for the iron has united with the oxygen in the air, and this union was so violent that heat and light were produced.

When a light is applied to a common candle, the light raises the temperature of the wick, the carbon and hydrogen which compose it, meeting with the oxygen of the air at this temperature, unite, and by their union more heat is evolved, and flame is the consequence, and this heat decomposes the tallow or wax of the candle, both of which are composed of oxygen, hydrogen, and carbon, and these again, uniting with the oxygen of the air, keep up the heat, and continue to produce flame—and so the candle burns. The flame of a candle may be divided into three parts—that near the wick, which is bluish; a second larger, which gives light, and is of a yellowish colour, enveloping the bluish centre; and a third, a thinner layer between the light-giving part and the air.





The centre part A, is where the gases evolved from the candle are collected, no burning takes place here. If a piece of paper be placed quickly in the flame, bringing it from above downwards, until it reaches the

middle of the blue part, but so as not to touch the wick, the paper in the blue part will not be burnt, showing that combustion is not going on there, but the part immediately surrounding it will be destroyed at once. If a flame could be cut through at the point indicated by the dotted line, three circles would be seen. A represents the blue centre, B the luminous envelope, and C the outer coat, which does not give much light, but which is intensely hot. When pure hydrogen and oxygen are burnt, very little light is given; but if into the flame a piece of charcoal be inserted, a bright light will be emitted; or, again, if a lump of lime be used instead of the charcoal, a strong brilliant light is produced, as is seen in the oxyhydrogen light, the solid charcoal or lime becomes heated to whiteness, and so gives light. This effect produced on solid bodies by heat, is called incandescence; and it is the solid particles in the flame which, becoming incandescent, cause light.

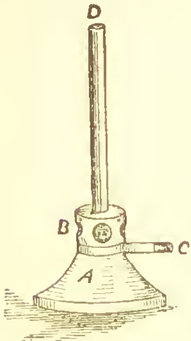
In the central part of the flame A the gases are collected, there, meeting with the oxygen of the air, when raised to a proper temperature, they unite with it, and here the union is principally of oxygen and hydrogen; small particles of solid carbon are given off, and these becoming incandescent in the second part of the flame B, through the heat produced by the union of oxygen and hydrogen, another combination takes place between carbon and oxygen; but the carbon here only unites with a certain quantity of oxygen. It has the power of uniting with more; it, however, unites with the smaller proportion because the carbon is in excess of the oxygen. When, however, this gas comes to the top of the part B of the flame, it meets with more oxygen, as it comes in

contact with the air, and so a second compound of oxygen and carbon is formed, when the carbon is completely burnt, that is when it unites with as much oxygen as it can unite with, and so, as a larger amount of chemical combination takes place, greater heat is evolved, and as the solid carbon has nearly all united with oxygen to form a gas, less solid particles being in that part of the flame, viz. the outer envelope C, it is less luminous.

That the central part of the flame A contains unburnt gases may be shown by placing a small tube with one end in the part A, and, by applying a light to the other end, the gases will pass through the tube, and burn at its extremity where the light is applied.

From what has been said it will be seen that when there is less oxygen and more solid carbon the flame gives more light, but when the oxygen is in excess, the carbon unites with it to form a gas, and hence there is less light and more heat. Now if a mixture of oxygen and coal gas be burnt (the coal gas is nearly the same in its constitution as the gas formed by the decomposition of the tallow or wax of a candle), there will be less light and more heat,—so that where greater heat is required such a mixture is employed. In chemical laboratories use is made of this source of heat. Coal gas and air are burnt together in a lamp called a Bunsen burner,—this lamp is so con-

structed that atmospheric air and coal gas mix together and escape together from a tube, and are then burnt. The lamp is made as follows:—



There is first a stand, A, and on this is fixed a hollow ring of iron, B, and in B four holes are made for admission of air; a pipe, C, enters B below the holes, and this conducts the coal gas into the lamp; the air and coal gas mix together and escape at D, where they are lighted. The flame of this lamp gives very little light but a large amount of heat. You will often notice outside butchers' or greengrocers' shops gas lamps which are not protected by glasses; on a calm night these give a steady uniform light, but when the wind blows on them the light decreases, and the flame is scarcely visible; this is owing to a larger supply of oxygen entering the flame, and so causing a more perfect combustion of the carbon.

From what has been said it will be easily understood that the different parts of a blow-pipe flame produce different effects on certain bodies when heated in them. As in the inner envelope carbon is in excess; if a metallic oxide be heated in it, its oxygen will be taken from it, and the metal will be left, and from the fact of the metallic oxide being reduced to the metallic state, the inner part of the flame is called the reducing flame; but if a metal which can be oxidised under such circumstances is heated just beyond the point of the flame where oxygen is in excess, it becomes converted into oxide, and that part of the flame is called the oxidising

flame. An experiment to show this is easily tried by heating a piece of ordinary lead glass tubing in the inner part of the blow-pipe flame, the oxide of lead in the glass is reduced, and black metallic lead is left, so that the glass at the heated part is black; if the blackened part be again heated—at the apex of the flame—the blackness will disappear, and the glass will become clear again owing to the oxidation of the lead.

As combustion is the result of the chemical combination of bodies, it is clear that bodies which are capable of uniting can do so in different ways: ordinarily hydrogen burns in oxygen, but oxygen can burn in an atmosphere of hydrogen. When speaking of oxygen we said that it could not burn in air,—this is because under ordinary circumstances there is nothing in air for it to combine with, and therefore there cannot be any combustion of it; but if a jet of oxygen be passed into a vessel full of hydrogen it will burn with a pale blue flame. The way to perform this experiment is to fill a large glass bottle with hydrogen, and to light the hydrogen at the mouth of this bottle, then to pass into it through the flame a tube connected with a gasholder containing oxygen, and through which a stream of oxygen is passing. The oxygen as it comes in contact with the flame at the mouth of the bottle will take fire, and if passed into the bottle will continue to burn as long as any hydrogen is uncombined,—the product of combustion is water, just the same as when hydrogen is burnt in oxygen. Combustion can take place where no oxygen is present, for as it is simply chemical union, or an effect of it, any sub-

stances which can unite can produce it; thus if iodine and phosphorus be put together they almost immediately take fire, phosphoric iodide being formed. Again, if turpentine in small quantities be dropped into chlorine gas, flame is immediately produced. Antimony also catches fire in chlorine.

Questions—Combustion.

1. What do chemists mean by combustion?

2. When hydric sulphate is poured into water, what physical effect is observed, and to what is it to be attributed?

3. Describe the different parts of a candle flame, and the chemical actions going on in a flame.

4. To what element does a flame owe its light? What is the chemical principle of the Bunsen lamp?

5. What is meant by incandescence? Give examples.

6. Describe a process by which heat is evolved in the bodies of men and of the higher animals.

7. What is meant by the oxidising flame of the blow-pipe, and in what part of the flame are metallic oxides reduced?

8. If the holes in the bottom ring of a Bunsen burner were stopped up, what effect would it have on the flame?

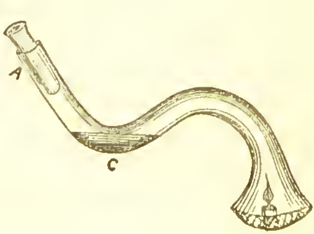
9. What effect has a strong current of air on an ordinary gas flame, supposing it be not sufficiently strong to extinguish the flame?

CHAPTER VII.

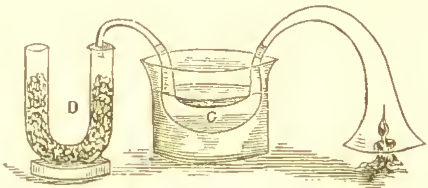
WATER—ITS COMPOSITION.

When oxygen and hydrogen burn slowly, as in the combustion of a lamp or candle, or when, being mixed together in proper proportions, they burn suddenly on the application of a light, accompanied by a loud explosion, or when they unite directly in any other way, water is formed. If a glass vessel be held over a burning candle or gas flame, it will become covered with water, and the vessel will appear as if it had been held over the mouth of a kettle; for the water formed by the burning of oxygen and hydrogen is, at first, from its high temperature, in the form of steam, but it becomes condensed on the sides of the cold vessel. An apparatus can be constructed in which all the water formed by the burning of a paraffin candle can be collected, and if the water be weighed, we shall find that it contains all the hydrogen which was in the substance of the candle, together with oxygen which that hydrogen has taken from the air to form water.

A is a bent glass tube with its end expanded into a bell shape. A candle is fixed on a piece of iron wire gauze, and is so fastened to the end of the bell expansion



that the flame of the candle is within the tube; the air necessary for burning the candle passes through the wire gauze, and the water formed collects in the bend of the tube C. It is necessary to keep the tube cool, and for this purpose a narrow long glass vessel, called a test tube, is filled with cold water and placed in it at A.



A further arrangement can be made to collect the other gas (carbonic acid) which is formed by the union of carbon and oxygen. If the tube be bent so that its part C is immersed in a vessel of water, this will cause the steam formed by the combustion of the candle to be condensed, and the carbonic acid will pass on into the tube D, which contains potassic hydrate. The potassic hydrate will absorb all the carbonic acid, and its increase of weight will show the quantity of it formed by the combustion of the candle; and from this it is easy to calculate the quantity of carbon which the consumed portion of the candle contained; so that all the products of the combustion of a candle, or any other substance which contains carbon and hydrogen, can be collected and weighed; and such an experiment as this is a good illustration of the fact that when a substance is burnt, nothing is lost, but its constituents exist in the same *quantities* as they did at first, but in different *combinations*. *Pure* water contains, then, nothing but oxygen and hydrogen.

We have now to consider in what proportions these two elements exist in water. If we weigh out 18 grammes of water, and convert the water into steam, and pass it over red-hot iron, as in a former experiment—the iron also being weighed—after all the steam has been carefully passed over it, we shall find that the iron has increased in weight 16 grammes, and that, as you know, is owing to its having taken the oxygen from the steam; and if the hydrogen, which passes on, be collected and weighed, it will be found to weigh 2 grammes; therefore in 18 grammes of water there are 16 grammes of oxygen and 2 grammes of hydrogen. Now, if you desire to find out how much oxygen there is in 50 grammes of water, you can do so by working out the following statement:—

Grms. of water.		Grms. of oxygen.		Grms. of water.
18	:	16	: :	50

or, $18 : 50 :: 16 : x$

and by a similar statement the amount of hydrogen in any given quantity of water can be determined, *i.e.*

Grms. of water.		Grms. of hydrogen.		Grms. of water.
18	:	2	: :	50

or, $18 : 50 :: 2 : x$

You will perhaps remember that when atmospheric air is passed over red-hot copper, finely divided, the oxygen of the air unites with the copper, forming black cupric oxide, and the nitrogen passes over free. If now we take a glass tube containing some of this black cupric oxide, and if we make it red-hot and pass over it pure hydrogen—say 2 grammes weight—the hydrogen will

take the oxygen from the cupric oxide, and, as when oxygen and hydrogen unite directly they always form water, so the product of this action will be water, which we can collect in a suitable apparatus and if we weigh the water produced, we shall find its weight to be 18 grammes. If there had been 20 grammes of oxygen in the cupric oxide, 2 grammes of hydrogen could not take from it more than 16 grammes of oxygen. Now suppose we have a substance which can yield up 50 grammes of oxygen to hydrogen, and we wish to know how much hydrogen we must take to make all the oxygen into water, you can find the quantity by working out this statement :—

	Grms. of oxygen.		Grms. of hydrogen.		Grms. of oxygen.		
	16	:	2	::	50		
or,	16	:	50	::	2	:	x

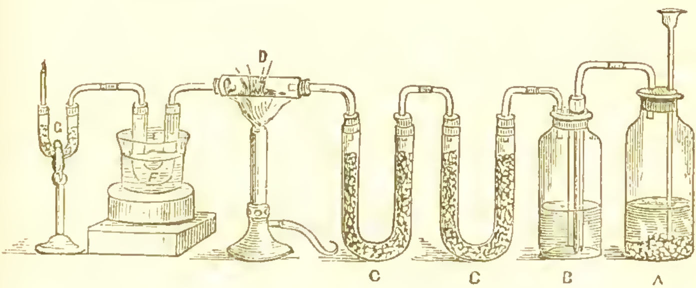
and in a similar manner the quantity of oxygen necessary to form water with 50 grammes of hydrogen can be determined :—

	Grms. of hydrogen.		Grms. of oxygen.		Grms. of hydrogen.		
	2	:	16	::	50		
or,	2	:	50	::	16	:	x

In the first of these experiments, we have found out the proportions of oxygen and hydrogen, by weight, in water, by taking the water to pieces, so to speak, and this is called *analysis*, which is derived from two Greek words—ἀνα (*ana*, back), and λυω (*luo*, I loosen).

In the second we have formed water by making oxygen and hydrogen unite, and by taking a given weight of hydrogen we have determined the exact proportion by weight in which they do unite to form water,

and so we have found out the composition of water by weight by a method which is called synthesis—which word means a putting together, and is derived from two Greek words—συν (*sun*, together), and τιθημι (*tithemi*, I place). In performing this last experiment with cupric oxide, it is necessary that the hydrogen should be absolutely pure and free from moisture. To ensure its being so the gas is passed through bottles containing argentic nitrate and afterwards through tubes containing calcic chloride, the argentic nitrate takes from the hydrogen its gaseous impurities, which have been already described, viz. arsenic, sulphur, and perhaps phosphorus, and the calcic chloride takes from it all moisture.



A is the bottle in which the hydrogen is generated, B contains argentic nitrate in solution, C C are tubes containing calcic chloride, D is the tube containing the cupric oxide, E is a vessel containing water in which is placed the receiver for the water formed by the reaction, and G is another tube containing calcic chloride to stop any water not condensed in F. If the tube D containing the cupric oxide be weighed before the operation, and if F and G be also weighed at the same time,

then the loss of weight of D after the operation will show the quantity of oxygen lost by the cupric oxide, and the increase of the weight of F and G will show the quantity of water formed, and therefore the quantity of hydrogen which has combined with oxygen to form it.

In this method of forming water it is usual to weigh the cupric oxide, not to measure the hydrogen, as that would be a more difficult operation: 79.5 grammes of cupric oxide contain 63.5 grammes of copper and 16 grammes of oxygen, and these 16 grammes of oxygen require 2 grammes of hydrogen to unite with them to form water—so that 79.5 grammes of cupric oxide will furnish 18 grammes of water. How much water can be obtained by passing hydrogen over 50 grammes of cupric oxide?

Grms. of cupric oxide.		Grms. of water.		Grms. of cup. oxide.
79.5	:	18	: :	50

or, 79.5 : 50 : : 18 : x

Again, how much cupric oxide must be taken to yield 20 grammes of water?

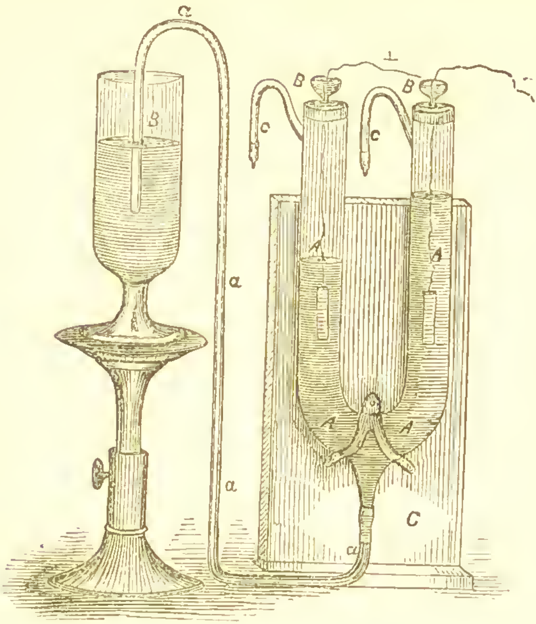
Grms. of water.		Grms. of cupric oxide.		Grms. of water.
18	:	79.5	: :	20

or, 18 : 20 : : 79.5 : x

When water is decomposed so that both its elements are free, they will both be in the state of gas; in this condition we can determine what *measure* or *volume* of oxygen and of hydrogen exist in a measure of water—water, you know, can exist in the state of steam, and in this state it is a compound gas. If the temperature

of our earth were sufficiently high, and we were so constituted as to bear that temperature, we should have no water, for it would all be converted into steam, and we should speak of it as a gas, just as we do of other gases. When, then, we speak of the volume of water in reference to the volumes of its constituents we mean its volume in the state of pure steam or gas. When the two wires or poles of a galvanic battery in action are placed in water acidulated with a little hydric sulphate the water is decomposed, bubbles of gas arising from each pole, and if two glass vessels of equal size and similar shape filled with water be placed with their mouths downwards, one over one wire-end or pole, and the other over the other pole, the bubbles will rise through the water and displace the water in the tubes. One tube will be found to fill with gas twice as fast as the other, so that when one is full the other is half full or nearly so. Oxygen dissolves more readily in water than hydrogen, therefore the volume of oxygen is rather less than half that of the hydrogen; this difference is, however, very trifling. This is called the electrolysis of water.

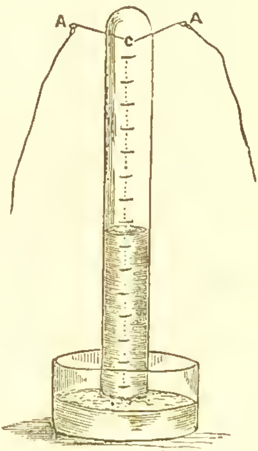
A very convenient apparatus for the decomposition of water by electrolysis is figured in the accompanying diagram. A, A, A, A is a U tube with a glass tube fixed into its curved part, and this communicates with a vessel of water, B, by the glass tubing, *a*, *a*, *a*, *a*. C is a stand for its support. + — show the wires from a galvanic battery,—passing into the legs of the U tube, and having at their extremities each a piece of platinum foil. As the gases are generated they force out the water into the vessel B, and their relative volumes can



be seen as shown in the diagram. There are small cups containing mercury, by which connection is easily established with the battery. The ends of the U tube, *a*, *a*, *a*, *a* are closed tightly with india-rubber corks. *c*,*c* are glass tubes fastened into the U tube; they are in the diagram represented as closed, but by them the gases can be led off and collected in a vessel or vessels, so that a supply may be obtained for any desired length of time.

If now we test these gases by the knowledge we have already acquired, we shall find that one,—that which fills its tube,—will burn, but the other, which half fills its tube will not burn; but if we put a light into it, that light will burn more brightly. We conclude then, that that which will burn is the hydro-

gen, and the other is the oxygen, and we see that one measures twice as much as the other, that is, the volume of one is double the volume of the other, therefore there is twice as much hydrogen as oxygen *by measure* in water, and this method of determining the volumes of the constituents of water by separating out these constituents is analytical. If 2 measures of hydrogen and 1 of oxygen be passed into a tube over mercury, and if an electric current be passed into the tube by the wires A A, there will be a spark formed at C which will heat the gases to a high temperature, and they will unite, forming water; and the mercury will rise in the tube until it nearly fills it as the steam formed will be condensed, which in the state of water will occupy about 1-2000th of the bulk of the original gases. Now, if more than 2 measures of hydrogen be taken and 1 of oxygen, the mercury will not rise to the top of the tube, as, whatever quantity of hydrogen there is above 2 measures (each equal to 1 measure of oxygen), will remain uncombined, and will occupy its own volume; that is, suppose 3 measures of hydrogen and 1 of oxygen be taken, after passing the spark, 1 measure of hydrogen will remain uncombined. Thus we see again that oxygen and hydrogen will only combine directly by measure, in the proportion of 2 volumes of hydrogen to 1 of oxygen; and we



have before seen that it will only combine by weight in the proportion of 16 grammes of oxygen and 2 grammes of hydrogen. This second example shows the measures of oxygen and hydrogen in water by the formation of water, that is, by synthesis.

The apparatus in which hydrogen and oxygen are made to unite by the electric spark over mercury is called a eudiometer. When the gases in the eudiometer are made to combine, if they be kept at the temperature at which water exists in the form of steam, the mercury will not rise to fill the tube after their combination, but will only rise one-third higher than it stood at first, because steam occupies two-thirds the volume, which the gases forming it occupy in the free or uncombined state.

If twice 11.2 litres of hydrogen be made to unite with 11.2 litres of oxygen, the steam formed will occupy the same volume which the hydrogen originally occupied, that is, twice 11.2 or 22.4 litres, and the weight of 22.4 litres of steam is 18 grammes, for 11.2 litres of hydrogen weigh 1 gramme; therefore 22.4 litres must weigh 2 grammes, and 11.2 litres of oxygen weigh 16 grammes; therefore 22.4 litres of steam are nine times as heavy as 22.4 litres of hydrogen, and 11.2 litres of steam, which weigh 9 grammes, are nine times as heavy as 11.2 litres of hydrogen, weighing 1 gramme, and therefore the density of steam is 9, hydrogen being 1. This method of calculating the density of compound gases is applicable in almost all cases (there are, however, one or two exceptions, which will be mentioned in their proper places).*

* The student is referred to page 2 in the Introduction.

An interesting and easy method of showing the energy with which oxygen and hydrogen combine is to mix them in the proportions in which they form water, and fill a strong glass bottle with the mixture—a soda-water bottle answers the purpose very well—and then to apply a light to the mouth of the bottle, a very loud report will be heard. It is always well in performing this experiment to wrap a thick cloth several times round the bottle, to prevent injury should the bottle burst. It is owing to the presence of oxygen and hydrogen that such terrible explosions take place when a light is brought into a room into which common coal gas has been allowed to escape. The oxygen in this case is supplied by the air. When, however, coal gas is exploded with oxygen, another product is formed besides water, viz. carbonic acid gas.

Questions—Water.

1. When oxygen and hydrogen unite directly, what body is produced?

2. How would you prove that water is produced by the combustion of a candle or of common coal gas in air?

3. Explain an experiment by which it can be shown that when a body is burnt nothing is lost.

4. 90 grammes of steam are passed over iron filings heated to redness, what will be the increase of weight of the iron, and how many grammes of hydrogen will be liberated?—*Ans.* The iron will increase 80 grammes in weight, and 10 grammes of hydrogen will be liberated.

5. Describe a method of showing the composition of pure water by analysis.

6. If 5 grammes of hydrogen be passed over heated cupric oxide, what weight of water will be formed?—*Ans.* 45 grammes. And what will be the volume of the steam?—*Ans.* 56 litres.

7. When water is decomposed into oxygen and hydrogen, how much space does the hydrogen occupy in comparison with the oxygen?

8. Steam is passed over red-hot iron filings and 35 litres of hydrogen are collected, what volume of steam was decomposed by the iron?—*Ans.* 35 litres.

9. 22.5 litres of oxygen are made to combine with a proportionate quantity of hydrogen, what will be the volume of the steam formed?—*Ans.* 45 litres.

10. It is required to form 108 grammes of water, what weight of cupric oxide must be taken and what volume of hydrogen?—*Ans.* 477 grammes of cupric oxide, 134.4 litres of hydrogen.

11. 40 cubic centimètres of oxygen are mixed with 50 cubic centimètres of hydrogen in a eudiometer and exploded, will any gas remain uncombined? If any, which of the two gases will it be, and how many cubic centimètres will it occupy?—*Ans.* Some oxygen will remain, and it will occupy 15 c. c.

12. 100 c. c. of hydrogen are exploded with 50 c. c. of oxygen, how many cubic centimètres will the steam occupy?—*Ans.* 100 c. c.

13. How many litres do 18 grammes of steam occupy, and how would you calculate the density of steam from this?—*Ans.* 22.4 litres; and as 1 gramme of hydrogen

occupies 11.2 litres, 9 grammes of steam will occupy 11.2 litres, therefore the density of steam is 9.

14. Describe the electrolysis of water. What are the results by volume, and to what apparent error is the method subject?

CHAPTER VIII.

WATER—ITS IMPURITIES.

Pure water contains nothing but oxygen and hydrogen. The water which we meet with in nature is always more or less impure, that is, it contains other substances besides oxygen and hydrogen. The purest water is rain water, but it absorbs gases on its way to the earth, and these of course render it, to the extent to which they are absorbed, impure. It is usual to divide waters into two classes, viz. fresh water and sea water. We will consider fresh water first. Water as it passes through the earth, dissolves out from the earth through which it passes whatever is soluble. The most common impurity is lime, and it is this which makes water hard; impurities of this kind are called inorganic or mineral. There are other impurities which are called organic, that is, they proceed from animal or vegetable substances, which have found their way into the water; and there are gases dissolved in water: these form the third class of impurities to which fresh water is liable.

There are different methods of getting rid of these: the first, or mineral impurities, may be got rid of

by distillation, that is, boiling the water and collecting and condensing the steam which passes over, and the mineral impurities are left behind; solid impurities may be separated by filtration through sand or porous stone; in laboratories filter-paper is used. When water is thick or muddy, and has a bad odour resulting from decaying organic matter, if it is filtered through animal charcoal, it becomes clear and sweet. Gaseous impurities can be driven off by boiling. Hard water contains lime, and its degree of hardness can be determined by Clarke's test, which consists in the addition to the water of a solution of soap in alcohol; the quantity of soap in the solution is known, and according to the quantity of soap used in producing the desired results, so the degree of hardness of the water is determined; each degree of hardness represents "one part by weight of carbonate of lime in 100,000 parts of water" —*Williamson.* It is well known that soap curdles when used with hard water; this is because the lime in the water forms a chemical compound, with one constituent of the soap, which compound is insoluble in water. When Clarke's solution is added to a measured quantity of water containing lime, on shaking them up in a bottle, no permanent froth will be formed, because the soap is decomposed by the lime; but on the addition of more soap-solution, when enough has been added to combine with the lime, then, on shaking the bottle strongly, a permanent froth or lather will remain, and the quantity of solution of soap used to produce this, tells the quantity of lime in the water, and therefore its degree of hardness.

There is a class of natural waters called mineral

waters, which are taken as medicine; some of these contain considerable quantities of saline matters, others contain gases, such as Harrogate water, which contains a gas called sulphuretted hydrogen, and there are chalybeate waters which contain iron. Seltzer water and natural effervescing waters are strongly impregnated with carbonic acid gas.

We are accustomed to call water colourless, but when it is viewed in large quantities, it has a bluish-green colour. If a long tube, about 15 feet long, closed at each end with glass, be filled with pure water, and if a strong light, such as the electric or oxy-hydrogen light be passed through it on to a white screen, it will be found to be coloured of a greenish tint. Water possesses a property most valuable to the chemist, as well as important in the economy of nature; it dissolves, or as it is commonly said, melts a vast number of substances. Common table salt, as we know, readily dissolves in water; and when the water has evaporated slowly, the salt will be found again at the bottom of the vessel, in pretty shapes, and with a sparkling appearance. The same thing happens if sugar be dissolved in water, and if the solution be evaporated to a thick consistency, and allowed to cool, solid masses are formed which are called crystals, and these crystals contain water, combined in a certain way with the sugar, and this water is called water of crystallisation. Many substances in the crystalline form contain water of crystallisation, such as epsom salts, glauber salt, alum, &c.; but others crystallise without having this water of crystallisation, such as potassic chlorate, argentic nitrate, &c. The water of

crystallisation can generally be driven off at a moderate heat; in some substances, such as calcic chloride, the heat required is greater.

When quick lime is moistened with water, it swells up, cracks, and falls to powder, and great heat is evolved. The quick lime takes into chemical combination a definite quantity of water, and this lime, so treated, is called slacked lime, and the water it has taken up is called water of hydration, and slacked lime is called chemically, calcic hydrate, quick lime being called calcic oxide.

Baryta is a substance similar to lime—it is an oxide of barium, and this also unites chemically with water, forming barytic hydrate. Soda and potash, which are oxides of sodium and potassium, also form hydrates. From these latter three the water of hydration cannot be expelled by heat, but from calcic hydrate it can be driven off at a red heat. There is a difference between water of crystallisation and water of hydration, but chemists cannot explain satisfactorily in what it consists, except that it seems to retain more of the properties which it had, as water, in the crystal than in the hydrate.

Sea water contains about 3.5 or 4 per cent. of saline matters, commonly called salts; the principal of these are sodic chloride (common salt), magnesic sulphate and chloride (to the presence of which the unpleasant bitter taste of sea water is owing), potassium salts, a small quantity of calcium salts, also bromides and iodides. Sea water is now distilled on board ship to be used as fresh water, but it tastes flat and insipid; if, however, it be passed through animal charcoal, it

dissolves the air condensed in the charcoal, and so, becoming aërated, is of a brisk and agreeable flavour.

The best test for lime in water is ammonic oxalate, which gives a white precipitate of calcic oxalate, *insoluble* in hydric acetate (acetic acid;) and this distinguishes it from calcic carbonate, which is soluble in hydric acetate. The tests for the other impurities of water will be given when the elements which they contain are treated of.

Questions—Water: Its Impurities.

1. What are the common impurities of water? How may they be got rid of?

2. In what way does water acquire its impurities?

3. To what is the hardness of water owing? How can the "degree of hardness" be determined?

4. What element does chalybeate water specially contain?

5. What is water of "hydration"? Name some hydrates, and the oxides from which they are formed. What is the effect of heat on sodic, barytic, and calcic hydrates respectively?

6. What is the percentage of solid matter contained in sea water? What are the substances found in it?

7. How would you prove the presence of lime in water?

CHAPTER IX.

ATMOSPHERIC AIR.

Atmospheric air, as has been stated, is a mixture of nitrogen and oxygen; in round numbers, the proportions in it of these gases are four-fifths of nitrogen and one of oxygen, or, more accurately, in 100 parts of air there are 20.9 parts of oxygen by volume to 79.1 parts of nitrogen, and by weight 23.1 parts of oxygen to 76.9 of nitrogen.

The nitrogen serves to dilute the oxygen so that it may be breathed continuously by animals; for if the air were all oxygen, it would destroy life by its violent action on the system. A piece of charcoal burns slowly in air; but if it be plunged when alight into oxygen the burning is much more rapid and is attended with brilliant light, which does not occur when it burns in air, and the consumption of the charcoal is much more rapid.

The oxygen and nitrogen in the air are not in a state of chemical combination; they are simply mixed together. This is proved in several ways. Water dissolves gases, some more readily than others; oxygen is more soluble in water than nitrogen. Now, when the oxygen and nitrogen of the air are dissolved in

water, it is found that they exist in it in different proportions to what they exist in air; this is easily proved by boiling the water in which air has been dissolved, and, as you know, boiling expels gases from water; when the gases are liberated from the water, on being analysed they yield not quite twice as much nitrogen as oxygen, whereas their relative proportions in air are 4 parts of nitrogen to 1 of oxygen; it is stated that their exact proportions when evolved from their solution in water is 1 part of oxygen to 1.87 of nitrogen. Now, if oxygen and nitrogen were chemically combined in air, they, being intimately united, must be dissolved in the proportions in which they exist in combination.

A substance called potassic pyrogallate has the power of absorbing free oxygen. When a measure of air is confined in a tube over mercury, if potassic pyrogallate be introduced into the tube, the mercury will gradually rise as the oxygen is absorbed, until about one-fifth of the space, originally occupied by the air, is filled with mercury. In performing this experiment, the pyrogallate and the potash should be put in separately, a small quantity of water being added to cause them to unite. Immediately the potassic pyrogallate comes in contact with the oxygen in the tube it changes from white to a brown colour, which gradually becomes deeper as more oxygen is absorbed. Neither pyrogallate nor potash *alone* absorbs oxygen. This absorption of oxygen from the air by potassic pyrogallate is also a proof that it exists free, and not in a state of chemical combination with the nitrogen, because if any gas which contains oxygen in chemical com-

bination be subjected to the same treatment as the air in the experiment just described, no change of colour takes place in the potassic pyrogallate, nor is the volume of the gas diminished.

There is a gas called nitric oxide, which will be described with the other oxides of nitrogen, which has this peculiarity, viz. that it unites with free oxygen to form a gas called nitrous acid. Nitric oxide is colourless, but nitrous acid is reddish brown. Nitric oxide does not take oxygen from substances which hold it in chemical combination. When nitric oxide mixes with atmospheric air it invariably becomes red, proving that the oxygen in the air with which it is united is free; and not chemically combined with the nitrogen.

These three proofs of the constitution of air are important, and should be remembered. Atmospheric air contains other gases besides the two mentioned. Carbon exists in all animal and vegetable substances, and as these are continually decaying, a gas called carbonic acid, which is formed by the union of carbon and oxygen, is being continually produced. It is, in fact, atmospheric oxygen acting on dead animal or vegetable matter which causes their decay. Animals, as we have seen, breathe in oxygen which unites with carbon in their lungs, and they breathe out this same gas,—carbonic acid. Fires, candles, gas lamps, all on burning give out carbonic acid. From these and other sources carbonic acid mixes in the air, and the quantity of it is said to be about 4 volumes in 10,000 of air. It is this gas, together with exhalations from the body, which causes the air in small and confined rooms to be

unwholesome; and as it, when present in large quantities, causes disease and sometimes death, it is necessary that such rooms should be well ventilated. "To keep an inhabited room in a state of good ventilation for one person 283 litres (10 cubic feet) of air should be removed every minute and replaced by fresh air."—*Williamson's "Chemistry for Students,"* § 46.

If water be exposed to the air in an open vessel it will in time disappear; this is because it is changed into vapour which escapes into the air. If the air be warm the water will evaporate into it more rapidly than when it is cold. From the sea, from rivers, ponds, and the moist surface of the earth, water is continually evaporating into the air, which must therefore contain a large quantity of water. The warmer the air, the greater the quantity of water vapour it will hold in suspension. Now, if air, which is charged with as much vapour as it can hold, be cooled down to a lower temperature, it will not be able to hold the same quantity of water vapour, and therefore water, as such, will be deposited. This is seen when a glass of very cold water is brought into a room in summer; the outside of the glass is soon covered with small drops of water or dew. This question may be asked you, How is water obtained from air? And the answer evidently is, by cooling that air and so causing the condensation of the water vapour.

The air ordinarily contains about half as much water vapour as it is able to hold in suspension. When the quantity is much less than this the air is too dry; this effect is produced when rooms are heated with closed stoves, such as an Arnott stove. In

this case it is necessary to have a vessel containing water in the room, which by its evaporation supplies moisture to the air. An excess of moisture is also unpleasant; as may be experienced by going into a laundry where clothes are drying.

Small quantities of a gas called ammonia are found in the air, which result from the decomposition of animal and vegetable matters containing hydrogen and nitrogen.

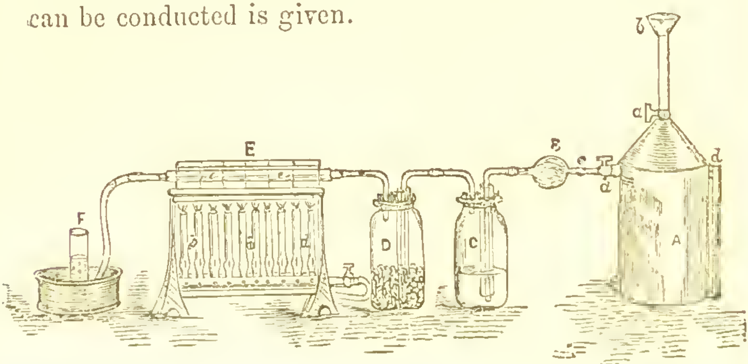
To recapitulate briefly what has been stated, and what it is absolutely necessary to remember; air contains about 1 part of oxygen and 4 parts of nitrogen, often a large quantity of water vapour, a small quantity of carbonic acid, and a still smaller quantity of ammonia.

There is also in air a variable quantity of a substance called ozone. This ozone is said to be oxygen existing in a different state, which is called "allotropic"; a further description of ozone in this place would lead us beyond the limits we have assigned to this book.

In order to obtain air pure, *i.e.* free from carbonic acid, ammonia, water vapour, and small particles of solid matter, it should be passed through vessels containing substances which will absorb these impurities. In making an analysis of air it should first be passed through cotton wool which will retain solid particles, then through caustic soda, which will combine with all the carbonic acid; there remain now no impurities, but water vapour and ammonia; these can be got rid of by passing it through a bottle containing lumps of pumice-stone moistened with hydric sulphate, the

principal constituents, oxygen and nitrogen, can be separated by passing the purified air over finely-divided copper heated to redness; the oxygen will unite with the copper, forming black cupric oxide, and the nitrogen passing on can be collected in a suitable vessel.

If a measured quantity of air be subjected to the above treatment, and if the bottle containing the sulphuric acid be weighed before the operation and after, its increase of weight will give the quantity of water contained in a given quantity of air; if the soda and the vessel containing it be weighed before and after the air has passed through it, the difference of weight will determine the quantity of carbonic acid; if the tube and the metallic copper contained in it be weighed before heating, and after the passage of the air, its increased weight will give the weight of oxygen, from which weight we already know how to calculate the volume, and the residual gas, nitrogen, can be collected and measured in a suitable vessel. This operation is that which is usually employed when it is desired to determine the quantities of the constituents of air; and a diagram of an apparatus in which it can be conducted is given.



A is a common gasholder filled with air, when the taps, *a a*, are turned and water is poured into the funnel *b*, the air is forced out through the tube *c*, d is a gauge by which it can be seen how much air is driven through this tube *c*, B is a glass bulb containing cotton wool, C a bottle containing lumps of fused caustic soda (sodic hydrate) moistened with water, D contains pumice-stone moistened with hydric sulphate, E is a furnace heated by a series of Bunsen burners *d d d*, through which the tube, *e e*, containing metallic copper passes (the square plates are represented as transparent in order to show the tube; they are in reality opaque, being made of fire-clay, and are used to retain the heat)—and F is the vessel in which the nitrogen is collected.*

Atmospheric air can also be analysed by the eudiometer. The use of this instrument was explained in the chapter on the composition of water. If 10 cubic centimètres of air be passed into a eudiometer standing over mercury, and to them be added 10 cubic centimètres of hydrogen, the volume over the mercury will be 20 cubic centimètres. If now an electric spark be passed through the mixed gases, all the oxygen in the air will unite with part of the hydrogen to form water, and the mercury will rise in the tube

* If it is desired to have the calculation *absolutely* correct, as the air will take up a very small quantity of moisture from the moistened sodic hydrate, it should be dried in a second bottle of pumice and hydric sulphate placed between the bottle of sodic hydrate and the tube containing the metallic copper. The moisture collected in this bottle need not of course be taken into account in the calculation.

and occupy the space which the gases, converted into water, occupied; and in this instance, that space will be 6 cubic centimètres, and 14 cubic centimètres of uncombined gases will remain.

Now, as water contains 1 volume of oxygen to 2 of hydrogen, the oxygen forms one-third of the volume of its constituents, therefore one-third of 6, or 2 cubic centimètres, will be the volume of oxygen contained in 10 cubic centimètres of air. In making a eudiometric analysis it is better to pass in more hydrogen than is necessary to combine with the oxygen; also, when the gases are exploded the eudiometer should be firmly pressed down on india-rubber, or some such elastic substance, to prevent the escape of the gases.

Animals require oxygen for their support; those which live on the earth take it from the air, but those which live in the water, as fishes, take it from the water; they do not take the oxygen which in union with hydrogen forms water, but that which, as you will remember, the water dissolves from the air. Trees, plants, and vegetables generally differ from animals in that they take in carbonic acid, whereas animals breathe in oxygen. Plants under the influence of sunlight take in carbonic acid, and by a very wonderful process they retain the carbon, which goes to form their structure, *i.e.* wood and its appendages, as well as other substances which they require for their own support and from which they form for our use starch, sugar, &c., and they give off oxygen—thus helping to keep up a a balance in nature between oxygen and carbonic acid, —for animals breathe out carbonic acid, which would, in time, cause their destruction, if it were not that

plants, requiring it for their sustenance, decompose it, and return the oxygen to the air whence it was originally taken. Here the operation is the reverse of what takes place in animals; in them, oxygen burns carbon and heat is evolved; in plants the carbonic acid which the animals have formed is unburnt, so to speak, and heat is stored up, for this unburning is a chemical decomposition; and when decomposition (that is, a separation of the substances, elementary or compound, which make up a body) takes place heat disappears, is absorbed, or becomes latent (as this fact is usually expressed). We may compare plants to a laboratory, in which substances usually called inorganic, or mineral, are converted into materials suited for the support of animal life. The grass which feeds and fattens cattle has, in the form of carbonic acid, ammonia, and water, taken in those elements which contribute most largely to the formation of flesh and fat, and in its leaves has changed them into substances which can be eaten and digested by animals, who in turn give back to the grass carbon, in the form of carbonic acid, and in their excreta, as manure, other substances which it requires for its nourishment.

The density of atmospheric air as compared with hydrogen is 14.4, for 11.2 litres weigh 14.4 grammes, and 11.2 litres of hydrogen weigh 1 gramme, and this is arrived at in the following manner:—11.2 litres of nitrogen weigh 14 grammes, and 11.2 litres of oxygen 16 grammes; and as there is four times as much nitrogen in air as there is oxygen, four times 14 grammes of nitrogen added to 16 grammes of oxygen

will give a volume equal to five times 11.2 litres, as, because there is no chemical combination, there will be no condensation of volume, and if this quantity be divided by 5, we shall get the weight of 11.2 litres of air.

14 multiplied by 4 gives 56, which added to 16 gives 72, and this divided by 5 gives 14.4.

Questions—Atmospheric Air.

1. Give the composition by weight and by volume of the gases in air.

2. What proofs are there that air is a mixture and not a chemical compound?

3. When air is acted upon by potassic pyrogallate, what gaseous body remains?

4. When nitric oxide is mixed with air what happens? If the oxygen of the air were in combination with the nitrogen, would nitric oxide have the same effect?

5. What other gaseous bodies are there in air, whence come they, and what purposes do they serve?

6. How can water and carbonic acid be *obtained* from air?

7. How much oxygen is there in 50 litres of air? Make this calculation from the exact proportion of oxygen in air.—*Ans.* 10.5 litres.

8. How much nitrogen is there in 500 litres of air? —*Ans.* 395 litres.

9. Give a method by which oxygen can be obtained in the free state from atmospheric air.

10. If air be passed through potassic or sodic hydrate, what change takes place in the air? and what in the sodic or potassic hydrate?

11. Explain how to make a eudiometric analysis of air.

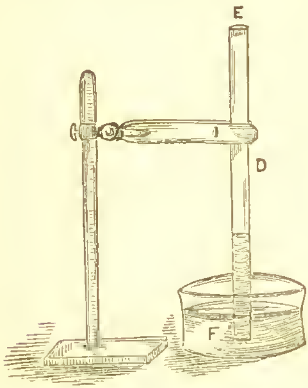
12. How much hydrogen is required to combine with all the oxygen in 20 cubic centimètres of air?—*Ans.* 8.4 cubic centimètres.

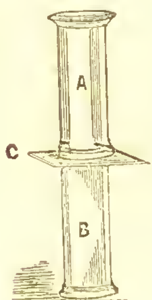
13. 5 volumes of air are mixed with 5 volumes of hydrogen in a eudiometer and exploded. What will be about the volume of the residual gases, and what their proportionate volume?—*Ans.* 7 volumes: 4 of nitrogen and 3 of hydrogen.

CHAPTER X.

DIFFUSION OF GASES.

Carbonic acid is heavier than atmospheric air; one would therefore naturally suppose that, as when two liquids are mixed together, as oil and water or mercury and water, the one being heavier than the other, the heavier sinks to the bottom, it would be the same with two gases of different densities, and that carbonic acid would, as it was formed, gradually displace atmospheric air from the earth's surface. This we know, however, is not the case, and the phenomenon by which the composition of the atmosphere is kept uniform, or nearly so, in all parts, is called diffusion. If the vessel A, filled with hydrogen, be inverted over





the vessel B filled with oxygen, and if the glass plate C, which separates the gases, be removed, in a short time both vessels will be filled with a mixture of oxygen and hydrogen. The oxygen has ascended to mingle with the hydrogen, and the hydrogen, though the lighter gas, has descended into the lower vessel. Again, if a glass tube, D, closed at one end by a piece of plaster-of-paris, E, be filled with hydrogen by upward displacement, when the plaster-of-paris is perfectly dry (the thumb being pressed over the plaster plug to prevent the escape of the gas during its collection) and be placed with its open end dipping into a vessel of water, F, after a short time the water will be observed to rise in the tube and afterwards to fall again to the level of the water in the vessel F. The rise of water in the tube is owing to the hydrogen passing through the plaster-of-paris plug more rapidly than the heavier atmospheric air passes in, and the pressure of the external air on the surface of the water in F, forces it up into the tube. After a time the hydrogen all diffuses out through the plug, and the equilibrium being established between the pressure inside and outside the tube, the water falls to its original level. It has been observed that gases diffuse into one another according to a fixed law. Hydrogen diffuses four times as quickly as oxygen; the density of hydrogen is 1, that of oxygen 16, the square root of 16 is 4, that of 1 is 1, so that the rapidity of the diffusion of oxygen compared with that of hydrogen is inversely as the square root of their densities, and this law is general. To find the rate of diffusion of any gas, with respect to hydrogen it is necessary to extract the square root of its density, and

then its rate of diffusion will be found to be as 1 to that square root. Diffusion is called a physical property. The unequal rates of the diffusion of the gases nitrogen and oxygen in atmospheric air afford additional proof that air is not a chemical compound.

Questions on Diffusion.

1. What is the rate of diffusion of hydrogen with respect to nitrous oxide?—*Ans.* 4.68.

2. What are the rates of diffusion of chlorine, nitrogen, and carbonic acid with respect to hydrogen?—*Ans.* Chlorine .168, nitrogen .267, carbonic acid .213.

3. Two vessels, each of 1 litre capacity, the one filled with hydrogen, the other with oxygen, are placed in a vertical position with their mouths together; after a short time a light is applied, what effect will be produced, and what will be the condition of the gases after this application?

4. If a vessel full of nitric oxide, and a vessel full of oxygen, be placed with their mouths together, what change will take place in the appearance of the gases? Explain the cause of this change.

CHAPTER XI.

CARBONIC ACID.

Carbon forms two compounds with oxygen. We shall first consider the most important of them, viz. carbonic acid. Whenever carbon is burnt in excess of air or oxygen, carbonic acid is formed. The diamond, on being burnt in oxygen, yields carbonic acid. Carbon, which has never been converted into vapour by intense heat, when heated, unites with oxygen very readily, and forms a new gaseous compound.

If carbonic acid be decomposed, 12 parts by weight of carbon and 32 of oxygen will be obtained. Suppose 12 grammes of carbon be heated in a current of oxygen gas—this can be done in a hard glass tube—and suppose the product (carbonic acid) to be carefully collected, it will be found to weigh 44 grammes, that is, 12 grammes of carbon and 32 of oxygen; and if it be got in the state of gas it will measure twice 11.2 litres, and therefore if we want to find its density with respect to hydrogen, by comparing its volume with an equal volume of hydrogen, as twice 11.2 litres of carbonic acid weigh 44 grammes, 11.2 litres will weigh 22 grammes; and as 11.2 litres of hydrogen weigh 1

gramme, the density of carbonic acid with respect to hydrogen will be 22. It should here be observed, as in the case of steam, so in the case of carbonic acid, when we take the weight of the constituents of either of them in grammes, and add them together, we get, in both cases, a quantity of the compound gas, which occupies twice 11.2 litres, and we find their densities with respect to hydrogen by halving that weight :—

$$\left.\begin{array}{l}\text{16 grammes}\\ \text{of oxygen, which}\\ \text{measure 11.2}\\ \text{litres}\end{array}\right\} + \left\{\begin{array}{l}\text{2 grammes of}\\ \text{hydrogen, which}\\ \text{measure 2(11.2)}\\ \text{litres}\end{array}\right\} = \left\{\begin{array}{l}\text{18 grammes}\\ \text{of steam, which}\\ \text{measure 2(11.2)}\\ \text{litres}\end{array}\right.$$

$$\therefore \frac{18}{2} = \text{9 grammes of steam, which occupy 11.2}$$

$$\therefore \text{9 is the density of steam.}$$

$$\left.\begin{array}{l}\text{12 grammes of}\\ \text{carbon}\end{array}\right\} + \left\{\begin{array}{l}\text{Twice 16 grs.}\\ \text{of oxygen, which}\\ \text{measure 2(11.2)}\\ \text{litres}\end{array}\right\} = \left\{\begin{array}{l}\text{44 grammes}\\ \text{of carbonic acid,}\\ \text{which measure}\\ \text{2(11.2) litres}\end{array}\right.$$

$$\therefore \frac{44}{2} = \text{22 grammes of carbonic acid, which measure 11.2}$$

$$\therefore \text{22 is the density of carbonic acid.}$$

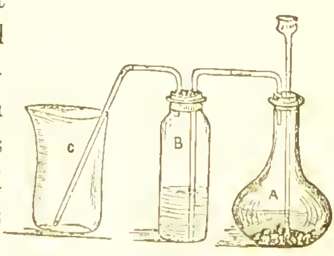
Carbonic acid is a colourless gas, 22 times as heavy as hydrogen; it has no smell, and turns moistened blue litmus paper slightly red, showing that it has acid properties (for one property of acid bodies is to change the colour of paper stained with a substance called litmus, and which is blue, to a red colour).

Animals cannot live in carbonic acid, nor can fire burn in it, as we have already seen, and therefore it is said not to support combustion; neither does it burn; but there are circumstances in which carbonic acid is decomposed, and then those bodies which decompose it

burn (that is, unite) with its oxygen, setting free its carbon.

There is a metal called potassium which, when heated, has power to decompose carbonic acid. If a small piece of this metal be placed in a glass flask containing carbonic acid, and if it be carefully heated, it will decompose that gas and will unite with its oxygen, forming white potassic oxide or potash, and black carbon will be deposited on the side of the flask.

Carbonic acid can be made by passing oxygen or air over red-hot charcoal; this is a real making of the gas: it is a synthesis; but for laboratory and other purposes it is obtained from one of its compounds. Common white marble is the best substance to employ. It is a combination of carbonic acid and lime (calcic oxide), which is called calcic carbonate. When this is mixed with an acid body called hydric chloride (a solution of the gas which is composed of hydrogen and chlorine being taken) a decomposition takes place, and three new substances are formed, viz. water, calcic chloride, and carbonic acid; the calcic chloride remains dissolved, and carbonic acid passes off, and being heavier than atmospheric air can be collected by what is called downward displacement, that is, the gas can be conducted by a tube from the generating apparatus to the bottom of a receiving vessel, and as it is heavier than air it will fall to the bottom just as water does when poured into



a glass, and so the vessel will become filled with

carbonic acid, and the air will be displaced. A is the vessel in which the marble and hydric chloride are placed together with a small quantity of water, B is a wash-bottle containing water, and is used to wash the gas, and retain any impurities which might come over with the spray caused by the effervescence, C is a beaker glass in which the gas is collected by downward displacement.

Carbonic acid can be collected over water, but the plan is not a good one, as water dissolves a very considerable quantity of the gas. As we have seen, carbonic acid exists in the air, in which its presence can be very easily detected.

Lime (calcic oxide) is slightly soluble in water, that is, if you take some common quick lime, and put it in water, and after a time filter off the clear liquid, some of the lime will be found to be dissolved in the water. Now carbonic acid unites chemically with calcic oxide and other similar oxides, forming calcic carbonate and other carbonates. Calcic carbonate does not dissolve in pure water—it is a white powder; if therefore carbonic acid unites with lime dissolved in water, it will form with it calcic carbonate, which does not dissolve in water, and we shall get a white substance formed which makes the water gradually milky, and which after a time settles to the bottom of the vessel; this substance is chemically of the same composition as chalk or marble, so that, if a solution of lime be exposed in an open vessel to the air, its lime will in time be deposited or precipitated as carbonate, showing that the air contains carbonic acid; and if this carbonate of lime be mixed with hydric chloride, the

same changes will take place as when marble is mixed with that body (for marble and chalk are the same chemically), and you can collect the carbonic acid by itself, and in this way can get carbonic acid from the air. From what has been said it is clear that carbonic acid exists largely in nature in combination with other substances, as chalk, marble and other carbonates. It is also found in magnesian lime stone, which is a carbonate of lime and magnesia.

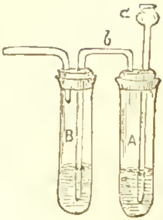
Neither hydrogen, oxygen, or nitrogen has been condensed to the liquid state; but carbonic acid can be liquefied by intense cold or heavy pressure. The usual method of liquefying it is by putting into a very strong iron vessel a carbonate with a liquid which will decompose it, and set the carbonic acid free. The carbonate usually taken is a carbonate of hydrogen and soda called hydro-sodic carbonate, and to this is added hydric sulphate, hydro-sodic sulphate being formed, and carbonic acid set free; when free the carbonic acid occupies a very much larger space than it did when in combination (you remember water in the state of steam occupies about 2,000 times its original volume). Now, if a quantity of hydro-sodic carbonate and hydric sulphate be taken sufficient to evolve a volume of the gas, about 36 times the measure of the iron vessel in which the condensation is to take place, and if these two substances be placed in it, in separate vessels, and after it has been tightly closed they be caused to mix, the carbonic acid will be evolved, and will press against the sides of the iron vessel, and the pressure will be so great that the gas will become liquid.

Or, carbonic acid may be liquefied by pumping it into a

strong iron receiver, until about 36 volumes have been compressed into the space of one,—at this point it will begin to liquefy: the liquid is clear and colourless. If the receiver be furnished with a stop-cock, and if this be opened, the gas will rush out with great violence, for the liquid will gradually return to the gaseous state, the pressure being removed; and in doing so, it will absorb so much heat from the particles of the liquid near to it, that they will become solid, and a snowlike mass of solid carbonic acid will be formed, the temperature of which is about —78°C.; this solid may be handled gently, but if it be pressed on the hand, it will cause blisters similar to these produced by intense heat.

Almost the greatest cold ever obtained was got by the vaporisation of a mixture of ether and solid carbonic acid in a vacuum. The pressure at which carbonic acid is liquefied is about 36 atmospheres, an atmosphere being a pressure equal to 760 mm. or 15 lbs. for the pressure of the air on the square inch is 15 lbs., or 760 mm.

Tests for carbonic acid: Carbonic acid being a very weak acid, is expelled from carbonates by hydric acetate (acetic acid), though hydric chloride is generally employed for its detection. To a solution of a carbonate, or to an insoluble carbonate suspended in water, hydric chloride is added, and the carbonic acid escapes with effervescence,—the gas should be led into lime water, and a white precipitate of calcic carbonate will be formed. The accompanying illustration shows a very convenient form of apparatus for performing this experiment.



A is a common test tube fitted with a cork, through which passes a funnel *a*. The liquid to be tested is put into this tube, and through the funnel a small quantity of hydric chloride is poured. As carbonic acid is soluble in water, if only a very small quantity be present, no effervescence will appear; it will then be necessary to warm the test tube. The gas passes through the tube *b* into the second test tube B, which contains lime water; in order to facilitate its passage, if the quantity be small, it is well to suck gently once or twice at the end of the exit tube, so as to draw the gas through the lime water, and the carbonic acid will convert the lime into carbonate, which will be precipitated.

Tests for carbonic acid in the gaseous state: It extinguishes a lighted match or candle—it has this property in common with nitrogen; it turns lime water white and milky, which nitrogen does not; and if carbonic acid gas be put in a tube over mercury and a sufficient quantity of a solution of caustic soda be introduced, the whole of the gas will be absorbed and the mercury will rise to fill the tube, except that part of it which is occupied by the soda solution.

Questions—Carbonic Acid.

1. What gas is formed when excess of air or oxygen is passed over red-hot charcoal?

2. How much carbonic acid is formed by the perfect combustion of a kilogramme of carbon?—*Ans.* 3.666 kilogrammes.

3. What volume does 264 grammes of carbonic acid occupy?—*Ans.* 134.4 litres.

4. How much oxygen can be obtained from 132 grammes of carbonic acid?—*Ans.* 96 grammes. And how much carbon?—*Ans.* 36 grammes.

5. What is the density of carbonic acid? Explain how you arrive at this density.

6. Give a general rule for finding the densities of compound gases and vapours.

7. What are the physical properties of carbonic acid? What are its chemical properties? State especially its effect on moistened blue litmus paper.

8. How can black carbon be obtained from carbonic acid? In this reaction what becomes of the oxygen?

9. Why cannot carbonic acid burn in air?

10. How is carbonic acid obtained for laboratory purposes? Explain the reaction which takes place, and name the substance which remains behind in the generating vessel.

11. How can the presence of carbonic acid in air be proved?

12. How can carbonic acid be obtained from atmospheric air?

13. What is the best method of collecting carbonic acid? Why is it better not to collect it over water?

14. How does carbonic acid exist in nature?

15. Explain the method of liquefying carbonic acid. By what mixture has very great cold been produced? Explain in what way this mixture produces cold.

16. What volume of atmospheric air is necessary to burn perfectly 30 grammes of carbon?—*Ans.* 280 litres.

17. A solution containing a carbonate is given—how would you prove the presence of carbonic acid?

18. A bottle contains a gas which extinguishes a lighted candle and renders lime water milky—what gas is it?

19. How would you distinguish between nitrogen and carbonic acid?

20. If a vessel contained a mixture of nitrogen and carbonic acid, how would you proceed to measure the quantity of nitrogen, and how to weigh the quantity of carbonic acid? If you wanted to determine the volume of carbonic acid, how would you do it?

CHAPTER XII.

CARBONIC OXIDE.

If carbonic acid gas be passed over hot charcoal and the product be collected, it will be found to differ in most respects from carbonic acid—this new gas is called carbonic oxide.

Carbonic oxide is colourless, tasteless, and has no acid properties like carbonic acid; it is also a violent poison, a very small quantity mixed with air causes death. Carbonic oxide burns with a pale blue flame, taking up more oxygen to form carbonic acid. The methods of obtaining this gas for laboratory and other uses will be given in a later part of this book.

If twice 11.2 litres of carbonic acid be passed over 12 grammes of red-hot carbon, the carbonic oxide formed will weigh 56 grammes, and if it be measured, its volume will be found to be four times 11.2 litres; therefore the carbonic acid, in taking up 12 grammes of carbon, has doubled its volume—and taking as our standard volume 11.2 litres, we say that 2 volumes of carbonic acid have become 4 volumes of carbonic oxide by uniting with 12 grammes of carbon, and these 4 volumes weigh 56 grammes (for the original 2 volumes, of 11.2 litres each, of carbonic acid weighed

44 grammes, and the carbon taken up by them weighs 12 grammes), and 2 volumes will weigh 28 grammes; and if we deduct from this 12 grammes, the quantity of carbon contained in 28 grammes of carbonic oxide, we shall have 16 grammes of oxygen left.

Now in 2 volumes of carbonic acid, each of 11.2 litres, we found that there were 12 grammes of carbon and 32 grammes of oxygen; therefore in 2 volumes of carbonic acid there is double as much oxygen as in 2 volumes of carbonic oxide, but there is the same quantity of carbon. Carbon therefore unites with oxygen in two proportions; that is, 12 parts of carbon unite with 16 parts of oxygen to form carbonic oxide, and with twice 16 parts of oxygen to form carbonic acid, and these combinations exist only in these proportions, for 12 grammes of carbon cannot unite with more than twice 16 grammes of oxygen, never mind how much oxygen be present; and 16 grammes of oxygen cannot unite with more than 12 grammes of carbon, never mind how much carbon be present. 16 grammes of oxygen occupy 11.2 litres, and these take up 12 grammes of carbon to form twice 11.2 litres of carbonic oxide; it is the carbon, therefore, in combination which doubles the volume of the gas. And as 11.2 litres of hydrogen weigh 1 gramme, carbon, with oxygen in the gaseous state, as carbonic oxide, occupies 2(11.2) litres, the oxygen in the uncombined state occupying 11.2 litres, therefore we may conclude that the density of carbon with respect to hydrogen is 12. At all events not less than 12 parts by weight of carbon can combine with hydrogen. (This will be better understood when the second part of this book has been read.)

Twice 11.2 litres of carbonic oxide weigh 28 grammes, therefore 11.2 litres weigh 14 grammes, so that the density of carbonic oxide with respect to hydrogen is 14.

It is often required to calculate the measure or volume of carbonic oxide which is formed when a given volume of carbonic acid is passed over red-hot charcoal. In every case the carbonic oxide is double the volume of the carbonic acid.

Having now considered carbonic acid and carbonic oxide and their formation, it will be easy to understand what takes place in the burning of an ordinary fire. When a fire is lighted air enters the bottom of the grate. The oxygen of the air being in excess of the carbon of the coals, carbonic acid is formed, the gas traverses a mass of hot coal (carbon) and of course becomes changed into carbonic oxide, which, issuing from the upper part of the fire, meets with fresh oxygen in the air with which it unites, forming a pale blue flame, and carbonic acid is produced; this pale blue flame which is seen flickering over the top of a clear fire is caused by burning carbonic oxide.

Water is also a product of combustion in all coal fires, the hydrogen in the coal forming water at the high temperature of the fire with the oxygen of the air. Carbonic oxide unites with potassic hydrate, if kept at a high temperature for some time; it forms a substance called potassic formiate. Carbonic oxide is dissolved by cuprous chloride in solution in hydric chloride.

Tests for carbonic oxide: It burns with a pale blue flame; it has no effect on lime water like carbonic acid, nor is it absorded by a solution of caustic potash under

ordinary circumstances; but if, say, 10 cubic centimètres of it be exploded in the eudiometer with 10 cubic centimètres of oxygen, 15 cubic centimètres of gas will remain, and if caustic potash be now introduced, 10 cubic centimètres will be absorbed, proving them to be carbonic acid, and the 5 c. c. which remain unabsorbed will be oxygen; for 5 c. c. of oxygen have united with the 10 c. c. of carbonic oxide without increasing its volume. You will remember that in 2(11.2) litres of carbonic acid there is twice as much oxygen as in 2(11.2) litres of carbonic oxide.

Questions—Carbonic Oxide.

1. How can carbonic oxide be obtained from carbonic acid?

2. State the physical and chemical properties of carbonic oxide. In what respects do its chemical properties differ from those of carbonic acid?

3. A volume of carbonic acid is passed over red-hot charcoal, what gas is formed, and what is its volume?

4. How much oxygen is required to convert 30 grammes of carbon into carbonic oxide?—*Ans.* 40 grammes.

5. 9 grammes of carbon are burnt in 9 grammes of oxygen, which element is in excess? What weight of it will be left?—*Ans.* Carbon 2.25 grammes.

6. 2 litres of pure air are passed over hot coke, what volume of carbonic oxide is formed? How much nitrogen is obtained?—*Ans.* .8 litres of carbonic oxide; 1.6 litres of nitrogen.

7. Explain what takes place in the burning of a common fire.

8. What is the density of carbonic oxide?

9. How do we arrive at the conclusion that less than 12 parts by weight of carbon cannot enter into chemical combination with hydrogen?

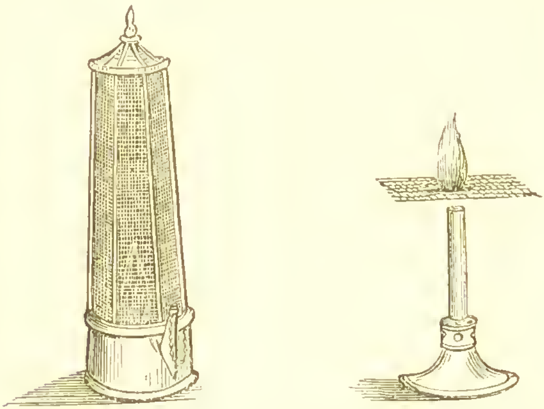
10. If 30 cubic centimètres of carbonic oxide be exploded in a eudiometer, with 15 cubic centimèters of oxygen, what will be the volume of the gases after explosion?—*Ans.* 30 cubic centimètres.

CHAPTER XIII.

COMPOUNDS OF CARBON AND HYDROGEN—MARSH GAS.

Marsh gas, or light carburetted hydrogen, is a colourless gas without taste or smell; it takes its name—marsh gas—from the fact that it is generated in marshes by the decomposition of vegetable matter containing carbon and hydrogen. It is also formed in coal mines, and is that which causes the fearful explosions which too often occur. It is called by miners *fire-damp*. It burns freely in air, with a pale but slightly luminous flame; but if it be mixed with air, and ignited, it explodes violently. The products of its combustion are water and carbonic acid. It is generated gradually in mines, and mixes with the air; when the mixture is such that the volume of marsh gas is one tenth the volume of the air, or half the volume of oxygen contained in the air, in any particular part of the mine, if it come in contact with a light it explodes. Carbonic acid and water vapour are formed, the water vapour is immediately condensed, and the carbonic acid remains—it is moist from containing water vapour, and is called *choke-damp*. It is as fatal to life, if not more so, than the explosion. No explosion can take place unless the combined gas and air are heated to a suffi-

ciently high temperature. It was to prevent these explosions that Sir H. Davy invented his safety-lamp, the principle of which is, that the surface of the lamp in contact with the outer air never gets sufficiently hot to kindle the gases outside the lamp. The lamp is surrounded with fine wire gauze, and the metal being a good conductor of heat, prevents the temperature from rising sufficiently high to ignite the gases outside it. The wire gauze employed should contain about 800 meshes in the square inch, and care must be taken that, when the lamp is in use, the wires do not rust and break, for if they do, ignition of the fire-damp takes place.



It is easy to try an experiment in illustration of this. Fine wire gauze is used in the laboratory; if a piece be held horizontally over the gas-burner, about two or three inches above it, and if a light be applied to the upper surface of the gauze, the gas will burn above but not below it, because the wire gauze conducts away the heat, so that the temperature never rises sufficiently high to ignite the gas below it.

On analysis marsh gas is found to contain 12 parts of carbon and 4 parts of hydrogen—that is, 12 grammes of carbon and 4 grammes of hydrogen form twice 11.2 litres of this gas; its density, therefore, with respect to hydrogen is $\frac{12+4}{2} = 8$. If it is desired to know how much oxygen is required to burn a measure of marsh gas—say 1 litre, we know that twice 11.2 litres of the gas will require 32 grammes of oxygen (which measure twice 11.2 litres) to burn its carbon (which weighs 12 grammes), so as to convert it into carbonic acid; and it will require the same quantity of oxygen to burn its hydrogen (which weighs 4 grammes), so as to convert it into water; therefore, twice 11.2 litres of marsh gas require four times 11.2 litres of oxygen for its complete combustion, that is, double its volume; and so 1 litre will require 2 litres of oxygen to burn it, or 10 litres of atmospheric air, and 1 litre of carbonic acid and 2 litres of water vapour will be formed.

Marsh gas is prepared by heating sodic acetate with soda lime, or with barytic hydrate; the hydrogen from the sodic hydrate (for soda lime contains sodic hydrate) or from the barytic hydrate unites with the hydrogen and part of the carbon of the sodic acetate, forming marsh gas. The remaining carbon and oxygen of the acetate form carbonic acid, and this unites with the sodium of the sodic acetate and the sodium and oxygen of the sodic hydrate to form sodic carbonate, the lime acts mechanically and remains for the most part as lime, though some may be converted into carbonate.

82 grammes of sodic acetate contain

	grms.
Sodium	23
Carbon	24
Hydrogen	3
Oxygen	32

40 grammes of sodic hydrate contain

	grms.
Sodium	23
Hydrogen	1
Oxygen	16

The sodic carbonate formed contains

	grms.
All the Sodium	46
Half the Carbon	12
All the Oxygen	48

The marsh gas formed contains

	grms.
Half the Carbon	12
All the Hydrogen	4

The changes which take place in this reaction will be more fully explained in the second part of this book. If marsh gas be passed through a tube containing pieces of pumice-stone, heated intensely, it is decomposed, carbon is deposited on the pumice, and hydrogen is set free, and the hydrogen is twice the volume of the marsh gas originally taken—for in twice 11.2 litres of marsh gas, there are four times 11.2 litres of hydrogen, and when the carbon is removed these are set free. Marsh gas has never been liquefied.

Questions—Marsh Gas.

1. How is marsh gas formed in nature?
2. What is fire-damp?
3. What is choke-damp? how does it destroy life?
4. Explain the principle of the Davy Safety Lamp. If a piece of wire gauze be held over a gas burner 2 inches from its orifice and the gas be lighted beneath the gauze, will the gas also burn above it?
5. How many volumes of hydrogen can be obtained

by decomposing 25 volumes of marsh gas?—*Ans.* 50 volumes.

6. How much atmospheric air is required to burn completely 5 litres of marsh gas? What volume of carbonic acid will be found, and what quantity of water vapour?—*Ans.* 50 litres of air, 5 litres of carbonic acid, 10 litres of water vapour.

7. What is the weight of 53 litres of marsh gas?—*Ans.* 37.857 grammes.

8. How much carbon can be obtained from 12 kilogrammes of marsh gas?—*Ans.* 9 kilogrammes.

9. Into a eudiometer containing 3 cubic centimètres of marsh gas, 8 cubic centimètres of oxygen are passed; the mixture is exploded at the ordinary temperature; what volume of gas will remain, and how much of that volume can be absorbed by potassic hydrate?—*Ans.* 5 cubic centimètres, of which 3 can be absorbed by potassic hydrate.

10. What are the products of combustion of marsh gas in oxygen?

11. Describe the preparation of marsh gas.

12. 100 litres of marsh gas are burnt in oxygen; what volume will their resulting bodies occupy in the gaseous state?—*Ans.* 300 litres.

13 How much atmospheric air is required for the complete combustion of 250 litres of marsh gas?—*Ans.* 2,500 litres.

CHAPTER XIV.

OLEFIANT GAS.

Olefiant gas (heavy carburetted hydrogen, or ethylene) is called olefiant because it forms, with chlorine gas, a heavy oily liquid. It is a colourless gas with a slightly sweet taste. It is formed when coal is heated in closed vessels; it is therefore a constituent of coal gas. It burns with a luminous flame, owing to the large quantity of carbon which it contains.

In olefiant gas there are 24 parts by weight of carbon, and 4 parts of hydrogen; it therefore has in it twice as much carbon as marsh gas. 24 grammes of carbon and 4 grammes of hydrogen make twice 11.2 litres of olefiant gas, and the sum of the weights of its constituent elements in grammes is 28, and 28 grammes occupy twice 11.2 litres; 11.2 litres contain 14 grammes; its density, therefore, with respect to hydrogen is 14.

Olefiant gas is made by decomposing alcohol, which contains the elements of olefiant gas and the elements of water. Hydric sulphate is used to effect this decomposition. You have before seen that hydric sulphate takes away the elements of water from sugar, leaving only carbon; but when it takes them away from alcohol, olefiant gas is set free. The hydric sulphate and alcohol

are mixed in the proportions by weight of 6 to 1, and are heated together in a retort, sand being added to moderate the action. The gas generally contains some impurities, such as the vapour of alcohol and a product of the action of hydric sulphate on alcohol, called ether; some hydric sulphate is also decomposed, and sulphurous acid passes over with the olefiant gas. The remaining oxygen from the decomposed hydric sulphate unites with some carbon of the alcohol, and carbonic acid is formed. If the gas be passed into a cooled receiver, the ether and alcohol vapour will be condensed. A solution of sodic hydrate will absorb the sulphurous and carbonic acids; and should any ether or alcohol still remain, they, together with water vapour, can be got rid of by passing it through the usual drying bottle containing pumice-stone moistened with hydric sulphate.

Olefiant gas is soluble in water, which dissolves about $\frac{1}{12}$th of its volume of it; it can also be condensed to a liquid by cold and pressure, but it has not been obtained in the solid state. Like marsh gas it can be decomposed by heat, but the heat required is less intense, so that when passed through a tube containing fragments of porcelain or pumice-stone heated to redness, carbon is deposited and marsh gas is formed; but if it be heated to an intense heat, hydrogen passes on, and this hydrogen, if the gas be completely decomposed, will be double the volume of the olefiant gas acted upon. Sulphuric acid and an acid called Nordhausen acid, both absorb olefiant gas freely, and in this way it can be separated from marsh gas. Hydric sulphate absorbs it, but slowly.

If equal measures of olefiant gas and chlorine be mixed together, they will, under the influence of light, unite to form the oily liquid called Dutch liquid, so that twice 11.2 litres of chlorine weighing 71 grammes, and twice 11.2 litres of olefiant gas weighing 28 grammes, form twice 11.2 litres of Dutch liquid in the state of vapour, weighing 99 grammes.

Olefiant gas, containing twice as much carbon as marsh gas, will require twice as much oxygen to burn the carbon, but as its hydrogen is the same, its combustion will require the same quantity of oxygen; therefore 1 litre of olefiant gas will require 2 litres of oxygen to burn the carbon, and 1 litre of oxygen to burn the hydrogen, that is, in all, 3 litres of oxygen, or about 15 litres of atmospheric air, and the products formed will be 2 litres of carbonic acid and 2 litres of water vapour—therefore, any measure of olefiant gas requires *three times its volume of oxygen* for its complete combustion.

Questions—Olefiant Gas.

1. How can olefiant gas be obtained from coal?

2. Explain the process of obtaining olefiant gas from alcohol. To what impurities is it liable? how can it be freed from them?

3. How many volumes or measures of hydrogen are there in 10 measures or volumes of olefiant gas?—*Ans.* 20.

4. How much oxygen is required for the perfect combustion of 50 litres of olefiant gas?—*Ans.* 150 litres.

5. Some olefiant gas is burnt with 6 litres of oxygen, what are the products of combustion, and what is their volume in the vapour state?—*Ans.* Carbonic acid and water vapour—4 litres of carbonic acid and 4 litres of water vapour.

6. What is the difference between the weight of a litre of marsh gas and a litre of olefiant gas?—*Ans.* .536 grammes.

7. If marsh gas and olefiant gas be mixed how can the marsh gas be obtained alone?

8. What volume of chlorine can combine with 5 litres of olefiant gas? and what is the name of the substance formed?—*Ans.* 5 litres.

9. Equal volumes of marsh gas and olefiant gas are mixed together, by their combustion they unite with 10 litres of oxygen; what weight of water will be formed, and what volume and weight of carbonic acid?—*Ans.* 6.428 grammes of water, 11.785 grammes of carbonic acid which measure 6 litres.

10. What is the density of olefiant gas? What other gaseous body has the same density?

11. What influence has light on a mixture of olefiant gas and chlorine?

CHAPTER XV.

COAL GAS.

Coal gas is made by the destructive distillation of coal. Coal is placed in an iron retort, with an opening for the escape of the gas; the retort is heated, and gas with other products are formed. Amongst these are tar, called coal tar, and ammoniacal liquor, that is, water holding ammonia in solution, and a crude coke is left behind. Coal tar contains other substances which can be separated out from it,—for example: Aniline, which has been used of late years in the manufacture of various beautiful dyes, is one of them. Creosote is also a product of coal tar. It is used for preserving meat from putrefaction. The peculiar smell and preservative properties of this substance are manifest in smoked meats, such as ham, bacon, reindeer tongues, &c.; it derives these properties from the presence in it of a body called phenilic hydrate, or commonly carbolic acid, which is now largely used as a disinfectant. Wood is also treated with creosote before being used as railway sleepers. Coal tar also contains several neutral oils, such as benzole, or benzine, as well as naphthaline; they contribute to the illuminating power of coal gas, and from benzole aniline is usually prepared.

Coal gas is a mixture; it consists mainly of hydrogen and the two gases we have just studied, viz. marsh gas and olefiant gas. Marsh gas forms a large part of the bulk of coal gas; the combustion of hydrogen and marsh gas causes the evolution of intense heat, but produces little light. Olefiant gas increases its luminosity; it contains about 5 per cent. of this gas. Although the hydrogen and marsh gas give but little light they hold in suspension naphthaline and benzole, which render them illuminating. An interesting experiment shows the change produced on hydrogen by these liquid hydrocarbons, and serves to illustrate the power which gases have of taking up other substances in the liquid state. If a current of hydrogen be passed through a bottle containing cotton-wool, or pumice-stone in small pieces, moistened with benzole, it will be found on ignition to give a brilliant light; whereas, before being passed through the benzole its flame was hardly visible. Various methods have been devised for rendering common coal gas more powerfully illuminating by adopting this treatment.

The coals found to give the best gas are cannel, and a coal called Boghead cannel; they yield much larger quantities than ordinary coal, Boghead cannel giving 15,000 cubic feet per ton, while ordinary coal yields about 9,000 feet.

The impurities to which coal gas is liable are ammonia and carbonic acid in small quantities. In two analyses given by Professor Roscoe—one of cannel gas, and the other of coal gas—the latter contained neither ammonia nor carbonic acid, and the former, in 100 volumes, only 3·07 of these substances

combined. The impurity which is of the greatest moment is sulphur, which exists in coal gas in two forms, as hydric sulphide (or sulphuretted hydrogen) and carbonic sulphide. Hydric sulphide may be got rid of by absorbing it with ferric hydrate, which is a compound of iron, oxygen, and hydrogen; the hydric sulphide consisting of sulphur and hydrogen is decomposed, the sulphur uniting with the iron to form ferric sulphide, while its hydrogen joins with the hydrogen and oxygen of the ferric hydrate to form water. Carbonic sulphide cannot, unfortunately, be removed by any means known of at present.

Carbonic oxide exists in coal gas, but it can hardly be called an impurity as it acts in a similar way to hydrogen and marsh gas, lending its aid in heating the flame, and evaporating the liquid hydrocarbons. Nitrogen is also an impurity of coal gas. The ammonia which is formed in gas manufacture being a valuable product, as it is a principal source from which sal ammoniac, an article of commerce, is obtained, is carefully separated out, and from the ease and completeness with which this can be done but small traces of it are found in the gas we burn. Its separation is effected by dilute hydric sulphate mixed with sawdust. A good notion of the composition of coal gas is given by an analysis of a Manchester coal gas by Professor Bunsen. In 100 volumes of the gas were found—

Hydrogen ...	45.58	Butylene ...	2.38
Marsh gas ...	34.90	Hydric sulphide	0.29
Carbonic oxide	6.64	Nitrogen ...	2.46
Olefiant gas ...	4.08	Carbonic acid ...	3.67

Questions—Coal Gas.

1. State some of the products formed in the destructive distillation of coal.

2. What gases are found in the state of mixture in coal gas?

3. How could you detect carbonic acid in coal gas?

4. How may coal gas be freed from hydric sulphide and ammonia?

5. Which gives the most luminous flame—olefiant or marsh gas? state why.

6. A common property of gases is to hold vapours in suspension; state how this affects the luminosity of coal gas.

CHAPTER XVI.

OXIDES OF NITROGEN.

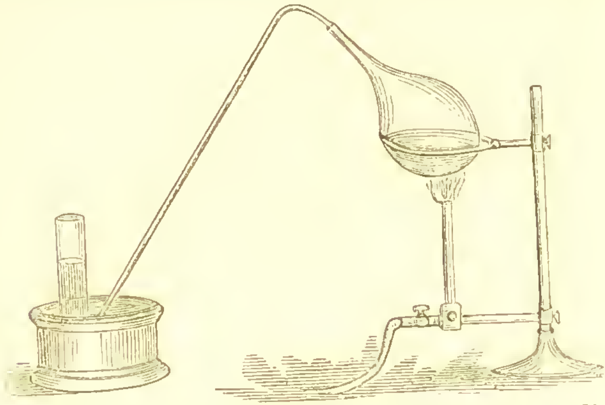
Nitrogen does not combine *directly* with oxygen. It will not burn in it. It exists in combination with oxygen in five different proportions, forming bodies which are called nitr*ous* oxide, nitr*ic* oxide, nitr*ous* acid, nitr*ic* peroxide, and nitr*ic* acid. When the terminations "ic" and "ous" are added they imply that more or less of a particular constituent exists in the compound. In the case before us that constituent is oxygen. Nitrous oxide contains the smallest quantity of oxygen, nitric oxide contains a larger quantity of the same element, and when a still larger quantity is present the prefix "per" is added, as nitric peroxide, which contains more oxygen than nitric oxide. It will be well also to notice the terms oxide and acid as they are *here* employed. In a future chapter the meaning of the term acid will also be more fully treated of. Those bodies which are called oxides, as carbonic oxide, do not usually unite with other bodies, elementary or compound, to form new substances, in the same way that acids do; but, as will be seen, those which are termed acids do unite with other bodies, which are called bases; as, for example, nitrous acid unites with

potassic oxide, forming potassic nitrite, and nitric acid unites with hydric oxide, or water, to form hydric nitrate. You will here perceive that two different terminations, "ite" and "ate," are used, expressive of the different quantities of oxygen contained in the different bodies formed by the union of an acid and base; the termination "ite" takes the place of "ous," and therefore indicates that a nitrite contains less oxygen than a nitrate, where the termination "ate" takes the place of "ic." In like manner, a sulphate contains more oxygen than a sulphite and a phosphate more than a phosphite.

Nitrous oxide is a colourless gas without smell and of a sweet taste. When inhaled it causes, in many cases, considerable excitement, which frequently manifests itself in laughter, hence it was called laughing gas. When breathed for some time it acts as an anæsthetic; that is, it renders the person breathing it insensible to pain. When used for this purpose it should not be mixed with atmospheric air.

Nitrous oxide is usually prepared from a compound of ammonia called ammonic nitrate, which contains the elements oxygen and nitrogen in the proportions to form nitrous oxide, together with the elements of water. When ammonic nitrate is heated it decomposes, and the nitrogen and some of the oxygen unite to make nitrous oxide, and the rest of the oxygen combines with the hydrogen to form water. The water becomes condensed, and the nitrous oxide passes on and can be collected over mercury or hot water; cold water should not be used for this purpose, as the gas dissolves in cold water to some extent. A drawing is here given of an

apparatus for the manufacture of nitrous oxide, which is also suitable for that of oxygen from potassic chlorate.



Nitrous oxide, like carbonic acid, can be made liquid by cold or by pressure; the pressure required is about 36 atmospheres. When the pressure is removed from this liquid it assumes the gaseous state with considerable violence, and some of it becomes solidified; its temperature when assuming the gaseous state is (— 88° C.) and it becomes solid at (— 101° C.) If liquid nitrous oxide, mixed with a liquid called carbonic disulphide, be exposed in a vacuum, most intense cold is produced, in fact the most intense which has ever yet been obtained, viz. (— 140° C.)

If a candle be blown out and plunged into a vessel of nitrous oxide while the wick is still red it will be rekindled and burn much more brightly than it does in air. Phosphorus burns almost as brightly in nitrous oxide as in oxygen; the gas is decomposed, the oxygen unites with the phosphorus, while the nitrogen is set free, and the nitrogen in its free state is found to measure exactly the same as the original nitrous oxide did.

Twice 11.2 litres of nitrous oxide weigh 44 grammes, but when the oxygen is taken away by burning phosphorus in it, twice 11.2 litres of nitrogen remain, the volume being unaltered; now 11.2 litres of nitrogen weigh 14 grammes, and twice 11.2 litres weigh 28 grammes, so that 44 grammes of nitrous oxide contain 28 grammes of nitrogen, and consequently 16 grammes of oxygen. And so we see that 2 volumes or measures of nitrogen and 1 of oxygen are condensed into 2 volumes in nitrous oxide gas. The density of nitrous oxide is 22, as the sum of the weights of its constituents in grammes is 44, and this weight occupies the volume twice 11.2 litres.

Nitric oxide is a colourless transparent gas, containing more oxygen than nitrous oxide, but it does not support combustion as readily. When a piece of lighted wood is placed in nitric oxide it goes out, unless the wood contains much resinous matter, so that its flame is very hot; for the gas must be decomposed before its oxygen can unite with other substances, and this requires a very high temperature.

Nitric oxide is usually prepared by decomposing hydric nitrate by copper. Hydric nitrate contains hydrogen, nitrogen, and oxygen; part of the oxygen and nitrogen unite with the copper, and form cupric nitrate; and another part of the nitrogen unites with oxygen in such proportions as to form nitric oxide, and the remaining oxygen unites with the hydrogen to form water; other metals, such as mercury and zinc, may be used instead of copper.

It is this gas which was mentioned when speaking of the constitution of the atmosphere; it readily

takes up free oxygen, forming nitrous acid, which is a red gas. If nitric oxide be analysed, that quantity of it which occupies twice 11.2 litres is found to contain 11.2 litres of oxygen and 11.2 litres of nitrogen; the weight therefore of twice 11.2 litres is 30 grammes (16 grammes of oxygen and 14 of nitrogen), and of course its density will be half that weight, that is, 15. Nitric oxide has never yet been condensed into a liquid.

Nitrous acid.—When nitric oxide unites with free oxygen nitrous acid is formed; it is, as has been stated before, a red gas, but it is usually made by acting on starch with hydric nitrate. Various combinations are formed, which remain in the retort, and nitrous acid is evolved. It is better made by the action of hydric nitrate on arsenious acid. It cannot be collected over water, as water decomposes it into nitric oxide and hydric nitrate. If, therefore, this red gas be shaken up with water in a bottle it loses its colour, the gas remaining in the bottle being nitric oxide, which is colourless, hydric nitrate being dissolved in the water.

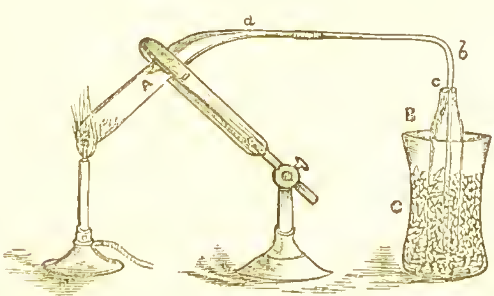
Nitrous acid is more easily condensed into a liquid than either nitrous oxide or carbonic acid, a pressure of about 4 atmospheres being sufficient for the purpose. If the gas as it is generated is passed into a tube surrounded by ice and salt, it will be collected in the tube as a deep blue liquid, and if the ends of the tube be sealed up, it can be kept for any length of time in that condition. (−18°C.) is the temperature at which nitrous acid liquefies.

Twice 11.2 litres of nitrous acid weigh 76 grammes, its density therefore is 38. The quantity of oxygen

contained in it is, by volume, three times 11.2 litres, and by weight 48 grammes, and the quantity of nitrogen twice 11.2 litres, and by weight 28 grammes; so that 5 volumes of the two gases are condensed into 2 volumes of nitrous acid; nitrous acid combines with bases to form nitrites.

Nitric peroxide is also a red gas, and can be condensed to a liquid, which, at a low temperature, crystallises in prismatic crystals. Liquid nitric peroxide is yellow at the ordinary temperature, but when cooled to (— 18°C.) it becomes colourless.

This gas is usually obtained by heating plumbic nitrate, which is a compound of lead, nitrogen and oxygen; the nitrogen in combination with part of the oxygen passes off as nitric peroxide, some of the oxygen remains with the lead, forming plumbic oxide, and the remainder of the oxygen passes off uncombined with any other substance. The density of this peroxide was for some time doubtful; it is now found that twice 11.2 litres measured at a low temperature weigh 92 grammes, its vapour density therefore is 46.



A very convenient apparatus for the condensation of nitric peroxide is given in the accompanying illustra-

tion. A is a piece of hard glass tube drawn out at its extremity, *a*, and sealed up at the opposite end; in this is placed the plumbic nitrate, and heat is applied; nitric peroxide and oxygen pass over into the vessel B, the mouth of which is drawn out and left just large enough to admit the delivery tube *b*, which passes nearly to its bottom; the glass beaker C contains a mixture of ice and salt; the gas is liquefied in B, and when sufficient has been obtained the end c is sealed with the blow-pipe. By substituting a flask or other generating vessel for the tube B, other gases which are condensible at the temperature, which can be obtained by a mixture of ice and salt, can be liquefied, such, for example, as nitrous acid or sulphurous acid.

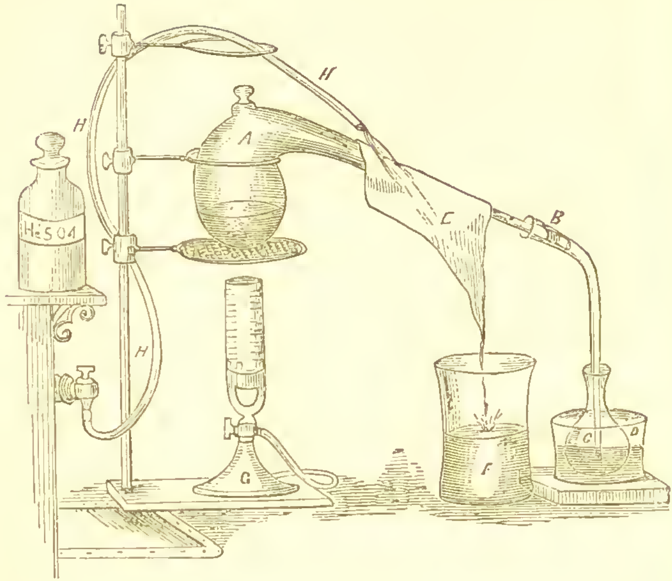
Nitric acid (sometimes called nitric anhydride and nitric pentoxide), is obtained in the form of transparent crystals by passing chlorine gas over argentic nitrate. Argentic nitrate consists of silver, oxygen and nitrogen. When this is melted, and chlorine gas is passed over it, argentic chloride is formed, and nitric acid and free oxygen pass over. The nitric acid can be caused to crystallise if passed through a tube placed in ice and salt, the crystals very easily decompose.

When nitric acid vapours are passed over finely-divided metallic copper at a high temperature, the oxygen unites with the copper, forming cupric oxide; and by taking a known quantity of the acid, the relative proportions of oxygen and nitrogen contained in it are found to be 80 parts of oxygen to 28 parts of nitrogen. Nitric acid combines with water (considerable heat being evolved) to form hydric nitrate.

Hydric nitrate is at ordinary temperatures a fuming liquid, colourless when pure, but usually yellow from the presence of some other oxides of nitrogen; it has strongly acid properties, and turns blue litmus paper to a bright red colour. When dropped on the skin it destroys it, turning it yellow; in a dilute state it is used to colour silk and woollen goods. It easily parts with some of its oxygen to bodies which combine with oxygen; and is therefore used as an oxidising agent. Hydric nitrate has been, and still is, called nitric acid, but reasons will be given in their proper place why this name is inaccurate. Hydric nitrate is also called aquafortis.

Hydric nitrate is made from potassic nitrate by replacing the potassium by hydrogen, and this is effected by taking about equal parts by weight of potassic nitrate and hydric sulphate, and heating them together in a retort; half the hydrogen of the hydric sulphate unites with the nitrogen and oxygen of the potassic nitrate forming hydric nitrate, which passes over and is condensed in a receiver, while the sulphur and oxygen, and half the hydrogen of the hydric sulphate together unite with the potassium, forming hydro-potassic sulphate.

A is a stoppered retort, into which the potassic nitrate and hydric sulphate are put; C is the receiver kept cool with water in the vessel D; B is an adapter, the use of which enables the flask C to be kept in a vertical position, E is a piece of filter paper, on which



water drops slowly from the india-rubber pipe H H H which is connected with a water tap ; the object of this is to keep the neck of the retort cool, and to condense the hydric nitrate as it distils over. F is a beaker to receive the water dropping from E, and G is an argand gas-burner for heating the mixture in the retort. A gentle heat should be employed.

If the hydro-potassic sulphate formed in this reaction be heated to a very high temperature with a second and equal weight of potassic nitrate, the remainder of the hydrogen leaves the hydro-potassic sulphate and combines with the nitrogen and oxygen of the potassic nitrate to form more hydric nitrate, and there remains behind potassic sulphate, the potassium of the second quantity of potassic nitrate takes the place of the hydro-

gen which has left it. 101 grammes of potassic nitrate contain 39 grammes of potassium, 14 grammes of nitrogen, and 48 grammes of oxygen; and 98 grammes of hydric sulphate contain 2 grammes of hydrogen, 32 grammes of sulphur, and 64 grammes of oxygen. When these weights of potassic nitrate and hydric sulphate are heated together, 1 gramme of hydrogen leaves the hydric sulphate, and, displacing 39 grammes of potassium, forms 63 grammes of hydric nitrate, and the 39 grammes of potassium take the place of the hydrogen which has left the hydric sulphate, forming 136 grammes of hydro-potassic sulphate; and when 136 grammes of hydro-potassic sulphate are heated very strongly with 101 grammes of potassic nitrate, the remaining quantity of hydrogen leaves the hydro-potassic sulphate, and, displacing the 39 grammes of potassium, forms another 63 grammes of hydric nitrate, and the 39 grammes of potassium take the place of the hydrogen, forming 174 grammes of potassic sulphate.

It is, therefore, easy to calculate the quantity of potassic nitrate which must be taken to yield any quantity of hydric nitrate. Say, for example, 50 grammes—

Grammes of potassic nitrate.		Grammes of hydric nitrate.		Grammes of potassic nitrate.
101	:	63	: :	50

Again, to find the quantity of potassic nitrate required to make 50 grammes of hydric nitrate—

Grammes of hydric nitrate.		Grammes of potassic nitrate.		Grammes of hydric nitrate.
63	:	101	: :	50

Also, if you wish to know the quantity of hydro

sulphate which will decompose 50 grammes of potassic nitrate so as to form hydric nitrate and hydro-potassic sulphate, you must work out this statement—

Grammes of potassic nitrate.		Grammes of hydric sulphate.		Grammes of potassic nitrate.
101	:	98	: :	50

Or, if you wish to convert the hydric sulphate into potassic sulphate, you must double the quantity of potassic nitrate, thus—

Grammes of potassic nitrate.		Grammes of hydric sulphate.		Grammes of potassic nitrate.
202	:	98	: :	50

You will observe in the changes we have been considering, that hydric nitrate is formed by the replacement of potassium in potassic nitrate by hydrogen from hydric sulphate. Thus, 1 gramme of hydrogen is capable of replacing 39 grammes of potassium in 101 grammes of potassic nitrate; but it requires twice 39 grammes of potassium to replace the 2 grammes of hydrogen in 98 grammes of hydric sulphate; but these 2 grammes of hydrogen are replaced in two successive operations, and *two* kinds of sulphates containing potassium are formed, but only *one* nitrate—that is, we have potassic nitrate, but we have no hydro-potassic nitrate, that is, no nitrate containing hydrogen and potassium, though we have two sulphates, viz. hydro-potassic sulphate, containing hydrogen and potassium, and potassic sulphate, containing potassium only.

If 56 grammes of potassic hydrate, which contain 1 gramme of hydrogen, 39 grammes of potassium, and 16

grammes of oxygen, be mixed with a quantity of hydric nitrate—say 100 grammes—an interchange will take place between potassium and hydrogen in the two compounds, hydrogen will leave the hydric nitrate, and potassium from potassic hydrate will take its place, and the hydrogen from the hydric nitrate will take the place of potassium in the potassic hydrate; the products formed will be potassic nitrate and hydric oxide, or water.

If we carefully boil this mixture, after a time all liquid will be driven off and we shall have a solid substance—potassic nitrate—left, and this substance will be found to weigh 101 grammes.

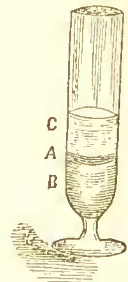
We originally took 56 grammes of potassic hydrate and 100 grammes of hydric nitrate, so that heat has driven off 55 grammes, consisting of water formed in the decomposition and of excess of hydric nitrate above what was necessary to combine with 39 grammes of potassium, to form potassic nitrate. This shows that, with excess of hydric nitrate present, no nitrate was formed, containing hydrogen and potassium, and that 39 grammes of potassium could only take the place of 1 gramme of hydrogen in 63 grammes hydric nitrate; for the solid residue after evaporation was 101 grammes, and this is the weight of potassic nitrate containing 39 grammes of potassium; for if excess of potassic hydrate over hydric nitrate be taken, the result as to the formation of potassic nitrate will be the same, only there will be free potash left after evaporating the water, but this can be easily separated from the potassic nitrate; so that if 63 grammes of hydric nitrate be mixed with an excess of potassic hydrate, 101

grammes only of potassic nitrate will be formed; therefore potassic hydrate and hydric nitrate can only react in the proportion of 56 to 63, and this proves that there is but one class of nitrates containing hydrogen, or some other metal in its place, and that there are none containing hydrogen and another metal, or a metal in more than one proportion to the other constituents of the nitrate. Hydric nitrate is also formed when nitric acid unites with water; this formation of the nitrate will be entered into more fully in the second part of this book.

Tests for a Nitrate.

The presence of a nitrate may be detected in any solution by first setting hydric nitrate free from any combination in which it exists. This is done by strong hydric sulphate, and if metallic copper be added the hydric nitrate is decomposed, part of its oxygen being taken away, as shown in the formation of nitric oxide, by the action of copper on hydric nitrate, and the nitric oxide shows itself by forming red fumes of nitrous acid when it comes in contact with atmospheric oxygen; if, however, the hydric nitrate be present in very small quantities the red fumes will not be perceived, but if a piece of paper moistened with starch and potassic iodide be placed above the liquid in the vessel containing the nitrate, the nitrous acid fumes will decompose the potassic iodide, and the iodine when set free will give a blue colour with the starch; the latter part of this test will be better understood after iodine has been studied.

There is also another good and very delicate test for hydric nitrate. If to a portion of the liquid to be tested an equal bulk of strong hydric sulphate be added, and the mixture be allowed to become *quite* cold, ferrous sulphate poured on the surface of this liquid will give a brown ring at its junction with it; the ring will be darker or lighter in proportion to the quantity of nitrate present: if there be much it will be formed at once; but if only a very minute trace it will appear after standing a short time. B is the liquid to be tested mixed with strong hydric sulphate, C is a solution of ferrous sulphate, and A is the brown ring at the junction of the two liquids.



These two reactions will, if obtained, afford positive proof of the presence of a nitrate, even if the quantity be extremely small. Nitric acid forms, with oxides of metals, compounds which are all soluble in water, therefore no substance gives a precipitate when mixed with it.

QUESTIONS ON HYDRIC NITRATE.*

1. Why is hydric nitrate often yellow? What is its action on the skin? For what purpose is it used in the arts?

2. Why is hydric nitrate used as an oxidising agent?

3. How is hydric nitrate usually prepared? What substance is left in the retort when the action takes place at a moderate temperature? When the sub-

* Questions on Oxides of Nitrogen given in Part II.

stances required to form hydric nitrate are mixed in proper proportions and subjected to a high temperature, what remains when the hydric nitrate has distilled over?

4. What weight of potassic nitrate and hydric sulphate must be taken to produce 693 grammes of hydric nitrate, if the action be carried on at a moderate temperature?—*Ans.* 1,111 grammes of potassic nitrate and 1,078 grammes of hydric sulphate.

5. What weight of hydrogen can be replaced by potassium in 343 grammes of hydric sulphate?—*Ans.* 7 grammes.

6. $50\frac{1}{2}$ grammes of potassic nitrate are heated with hydric sulphate, what quantity of hydric nitrate will be formed if the residue consists of potassic sulphate? —*Ans.* $31\frac{1}{2}$ grammes.

7. How much potassium replaces the hydrogen in 490 grammes of hydric sulphate when they are heated to a high temperature with potassic nitrate?—*Ans.* 390 grammes.

8. How much hydric sulphate is required to completely decompose 808 grammes of potassic nitrate?—*Ans.* 392 grammes.

9. How much potassic sulphate and how much hydric nitrate are produced when 40.4 grammes of potassic nitrate are heated with 19.6 grammes of hydric sulphate? —*Ans.* 34.8 grammes of potassic sulphate, 25.2 grammes of hydric nitrate.

10. Describe the two methods by which a nitrate may be discovered when in solution.

11. Why cannot the presence of a nitrate be detected by precipitation?

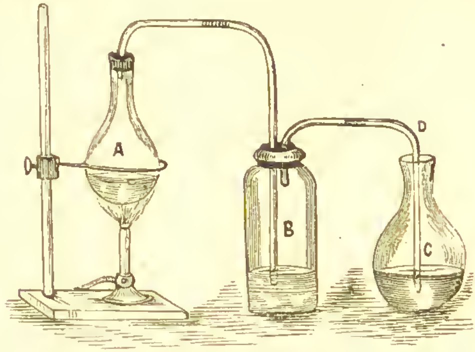
CHAPTER XVII.

AMMONIA.

This pungent and powerful-smelling gas is contained in salvolatile and hartshorn; it is transparent and colourless. It is formed by decaying animal and vegetable substances containing nitrogen and hydrogen. As we have seen, it is largely produced in the manufacture of coal gas. It was called ammonia from the fact that it was first obtained from a compound made by the Arabs in Libya, near a temple dedicated to Jupiter Ammon, and this compound was from this circumstance called sal ammoniac. The Arabs obtained it by heating the dung of camels.

Sal ammoniac, or ammonic chloride, contains chlorine, hydrogen, and ammonia; and when it is heated with lime (calcic oxide) it is decomposed in the following manner—the chlorine of the ammonic chloride unites with the calcium of the calcic oxide, forming calcic chloride, and the hydrogen of the ammonic chloride unites with the oxygen of the calcic oxide to form hydric oxide (water), and the ammonia is set free. 107 grammes of ammonic chloride, heated with 56 grammes of calcic oxide, give 111 grammes of calcic

chloride, 18 grammes of water and 34 grammes of ammonia gas.



The apparatus used for preparing ammonia consists of a flask, A, in which is put the ammonic chloride and calcic oxide, moistened with water so as to form a pasty mass. These are heated, and the ammonia gas passes over into a second vessel, B, containing a little water, which prevents impurities from passing into the vessel C, where the ammonia is finally received, and which may be of any convenient form; this vessel should contain distilled water, which absorbs the ammonia gas until it is completely saturated. It will be observed that at first the bubbles of gas rising from the conducting tube D are at their escape from its mouth large, but on passing through the water disappear before they arrive at its surface—this is owing to their absorption by the water. When bubbles of the gas escape at the surface of the water as large as they are on entering it from the conducting tube, it is a sign that the water is saturated; that is, it will absorb no more ammonia.

Ammonia is kept for use thus dissolved in water,

and we may regard it as combined with a proportionate quantity of water, forming a hydrate just as potassic oxide does; so that ammonic hydrate contains ammonium,* oxygen, and hydrogen, as potassic hydrate contains potassium, oxygen, and hydrogen. It only combines with a *proportionate* quantity of water; the whole of the water in an aqueous solution of ammonia is not so combined; any more than is all the water in an aqueous solution of potassic hydrate, for we can evaporate the water from potassic hydrate, and even heat it to a high temperature, and the mass when cold is a white solid—yet it is potassic hydrate, containing oxygen, hydrogen, and potassium in exactly the same relative proportions which it did when dissolved in water. But we cannot evaporate the ammonia solution in the same way, because ammonia, being a very volatile gas, is driven off as soon as heat is applied. It is necessary that this should be well understood, that many substances combine with water, but only in fixed proportions, and that the combination is a chemical not a mechanical one; but after this, if they are soluble in water, any addition of water can be expelled by heat, as we have seen in a former chapter, leaving the hydrate just as it was before that extra water was added.

If a volume of ammonia gas be collected over mercury in a eudiometer, and if a succession of electric sparks be passed through it, its volume will be found to become double what it was at first; this is because the

* Ammonium contains more hydrogen than ammonia; its composition will be explained in Part II.

gas has been decomposed. If now the gases in the eudiometer be analysed they will be found to be hydrogen and nitrogen, in the proportion of 3 volumes of hydrogen to 1 of nitrogen. Twice 11.2 litres of ammonia gas contain three times 11.2 litres of hydrogen which weigh 3 grammes, and 11.2 litres of nitrogen weighing 14 grammes, the weight therefore of twice 11.2 litres of ammonia gas is 17, and half this weight, 8.5, is the density of ammonia with respect to hydrogen. Always remember that the volume of ammonia gas is half the volume of its constituents—that is, if 3 volumes or measures, of any capacity, of hydrogen were made to unite with 1 volume or measure of nitrogen equal to one of the measures of hydrogen the resulting ammonia gas would measure just half 4 volumes, *i.e.* 2 volumes, each equal to 1 volume of the hydrogen employed; so that if we desire to know the volume of hydrogen and nitrogen in, say, 50 litres of ammonia gas, the 50 litres, if decomposed, will become 100 litres—one-fourth, *i.e.* 25 litres of which will be nitrogen, and the other three parts, *i.e.* 75 litres, will be hydrogen. Ammonia is also decomposed by passing it through a red-hot tube.

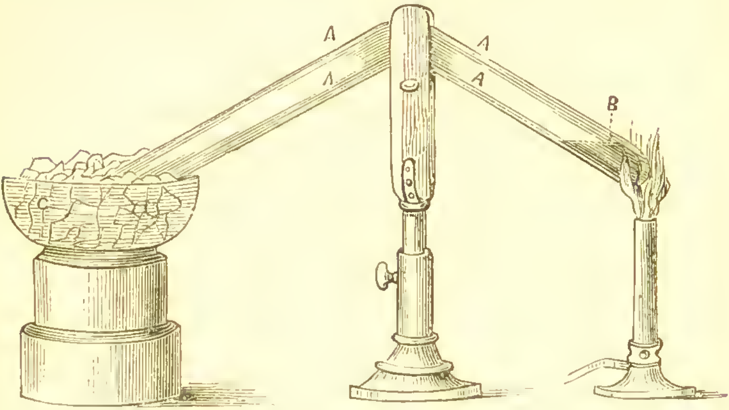
We have seen that acid bodies turn blue litmus paper red; ammonia and bodies like ammonia, which have alkaline properties, restore the blue colour to litmus paper, which has been reddened by acid bodies. Ammonic hydrate, like potassic hydrate, reacts with acid bodies, forming compounds which have no action at all on either blue or red litmus paper: such bodies are called neutral; for example, if hydric nitrate be added to a solution of ammonic hydrate, which con-

tains ammonium, oxygen, and hydrogen, the ammonium unites with the nitrogen and oxygen of the hydric nitrate forming ammonic nitrate, and the hydrogen of the nitrate unites with the oxygen and hydrogen of the ammonic hydrate to form hydric oxide (water), and the resulting ammonic nitrate has neither acid nor alkaline properties—it is neutral. It is also here seen that what has been called ammonium acts like potassium, which is a metal. Ammonium is by chemists considered to be a metal; it can take the place of hydrogen in hydrogen compounds, as in hydric sulphate or in hydric nitrate, and it behaves in all chemical respects as a metal. Other reasons for so regarding it, together with its constitution, will be given in the Second Part.

Ammonia gas can be condensed, like carbonic acid, but it requires less pressure, and the temperature at which it assumes the gaseous form, when liquefied, is higher than that at which liquid carbonic acid does so. The pressure necessary to liquefy ammonia gas is about 6.9 atmospheres; and the temperature at which it becomes gaseous after liquefaction is (— 38.5°C.) according to Regnault. By a cold of (— 75°C.) it is frozen to a white, transparent, crystalline solid, which is more dense than the liquid.

A convenient apparatus for liquefying ammonia is given in the accompanying illustration. AA is a very strong glass tube; in the part B is placed some dry argentic chloride, which has been exposed to the action of ammonia gas, and which has absorbed a large quantity of it. The tube is then sealed up and heat is applied to the part B. The end C of the tube AA is placed in a vessel containing ice and salt, and in this

part the ammonia which is driven off from the argentic chloride by heat is condensed to a liquid.



As ammonia is soluble in water, which will absorb above 1,000 times its own volume at 0°C., and 586 times its volume at 26°C., when it is required in the gaseous state, not as an aqueous solution, it is collected either over mercury or by upward displacement, like hydrogen; its density being 8.5, and that of air being 14.45, it is considerably lighter than air, and therefore can ascend.

In order to get dry ammonia gas it cannot be passed through sulphuric acid, or over calcic chloride, which is the general method for drying gases, as it forms compounds with them; but it should be passed over potassic hydrate or quick lime, both of which take up moisture readily.

Lighted bodies are extinguished when plunged into ammonic gas; it however burns, but not readily, in air, unless it be kept at a high temperature, by passing it through the upper part of the flame from a Bunsen burner. It burns more freely in oxygen than in air.

Tests for Ammonia.

Ammonia, even in small quanties, can be easily discovered. If the liquid containing it be heated with potassic hydrate, the ammonia is set free just as it is when ammonic chloride is heated with calcic oxide, the reaction being precisely similar, and the free ammonia can be detected by its smell, or if the quantity be small, by its forming dense white fumes of ammonic chloride on the introduction of a glass rod moistened with hydric chloride into the vessel containing it.

Nessler's test.—If mercuric iodide be dissolved in potassic iodide a clear colourless solution is formed; if after addition of sodic or potassic hydrate this be poured into a liquid containing ammonia (even the smallest trace) a brown precipitate, or if the quantity of ammonia be small, a yellow colour will be produced.

Questions on Ammonia.

1. Describe the physical properties of ammonia gas. How is it formed in nature? From what did it acquire its name? In the manufacture of what article of commerce is it largely produced?

2. How is ammonia prepared for laboratory use? Describe fully the reaction which takes place.

3. Of what does the liquid called ammonia, which is used in laboratories, consist?

4. 15.5 cubic centimètres are decomposed in a eudi-

ometer; what will be the volume of the gases after decomposition, and in what proportions will they exist?—*Ans.* 31 c. c.; nitrogen 7.75 c. c., and hydrogen 23.25 c. c.

5. How many volumes of oxygen are required to unite with the hydrogen in 4 volumes of ammonia gas?—*Ans.* 3 volumes.

6. At what temperature can ammonia gas be condensed to a liquid? Describe an apparatus for its condensation.

7. How can ammonia gas be obtained free from moisture?

8. Does ammonia gas burn in air or oxygen?

9. Describe the tests for ammonia.

10. How much ammonia can be obtained from 2,140 grammes of ammonic chloride? What quantity of water and what quantity of calcic chloride will be formed if calcic hydrate be employed for its decomposition?—*Ans.* 680 grammes; 720 grammes of water; 2,220 grammes of calcic chloride.

11. What is the action of potassic hydrate on a salt of ammonia?

12. What is the density of ammonia gas? Show how you arrive at it.

CHAPTER XVIII.

CHLORINE.

The next four elements which we are now about to consider are generally classed in a group, because they form similar compounds with other elements. They none of them exist in nature in the uncombined state. They all form compounds with hydrogen which have acid properties, and with the metals by displacement of that hydrogen, as well as by direct combination, and because these compounds, with the metals, are analogous to common salt, they have been called halogens, or salt producers. It will, however, be seen that this title belongs to an age in chemistry which has passed away. Neither of these elements combine directly with oxygen, *i.e.* they will not burn in it.

Chlorine was discovered by Scheele in 1774. It is a greenish yellow-coloured gas—hence its name, chlorine, from the Greek word χλωρὸς, *chloros*, green. It has a strong suffocating odour, and when inhaled in very small quantities irritates the air passages of the lungs, causing a painful sensation and coughing; and in larger quantities it destroys their lining membrane. Chlorine always occurs in nature in combination, forming compounds called chlorides. Common salt

contains chlorine and sodium; it is called sodic chloride, and is found in large quantities in certain mines, salt springs, and sea water.

Chlorine is a heavy gas, 11.2 litres of it weigh 35.5 grammes; its density, therefore, is 35.5 when compared with hydrogen.

Chlorine is somewhat soluble in cold water, which dissolves double its bulk; this prevents its being collected over cold water; warm water holds it in solution much less readily than cold, and is therefore better suited for this purpose. It cannot be collected over mercury, as it combines with it energetically, forming a chemical compound; it is best collected by downward displacement. Chlorine is easily liquefied At a pressure of 4 atmospheres it forms a yellow liquid, but it does not solidify at the lowest temperature known.

Chlorine combines very readily with metals. Powdered antimony, finely-divided iron or copper, and Dutch-metal leaf, when thrown into dry chlorine gas, catch fire immediately, owing to the union of chlorine with the metal used, and the resulting compound is called a chloride. Chlorine does not burn in air, but when a candle, turpentine, or bodies containing hydrogen and carbon, are placed in it alight, they continue to burn with a dull red flame, and a large quantity of smoke is formed; this is owing to the fact that the chlorine combines with the hydrogen of the burning substance, but not with the carbon (you will remember that oxygen combines with both, forming water and carbonic acid when similar bodies are burnt in it), which is deposited as soot. So great is the energy of combination, that when a piece of paper moistened with

turpentine is dropped into a vessel of chlorine, it takes fire immediately, burning with a smoky flame, and depositing carbon.

Chlorine has no effect on litmus paper when dry, but if the paper be moistened it immediately discharges the colour of the litmus. It also bleaches a solution of indigo. This bleaching effect of chlorine will be referred to again when speaking of the action of hydric hypochlorite.

Chlorine is used as a disinfecting agent, from the power which it has of destroying bad odours and decomposing hydric sulphide, commonly called sulphuretted hydrogen.

Chlorine is usually prepared from one of its compounds called hydric chloride, which contains it united with hydrogen. This body has powerfully acid properties: it turns blue litmus paper to a bright red, and is commonly called hydrochloric acid. Many years ago it was called by chemists muriatic acid, and by common people then, as well as now, spirits of salt, but its correct name is hydric chloride. When hydric chloride is heated with a substance which can take away its hydrogen, chlorine is of course set free. The substance used for this purpose is black oxide of manganese (properly called manganic binoxide), which contains the metal manganese and oxygen. The oxygen of the manganese takes the hydrogen of the hydric chloride and forms water, setting its chlorine free, half of which at the moment of its liberation unites with the manganese, forming manganic chloride, and the other half escapes, and can be collected as free chlorine. 146 grammes of hydric chloride require 87 grammes of

manganic binoxide for their decomposition; that is, 146 parts, whatever be the weight of a part adopted—whether grammes, grains, pounds, or ounces—require 87 parts of the same weight each for their perfect decomposition.

Following the simple plan used before, the quantity of manganic binoxide required to decompose a given weight of hydric chloride can be found by the following statement—

Grms. of hydric chloride.		Grms. of manganic binoxide.		
146	:	87	: :	given quantity of hydric chloride.

Now 146 grammes of hydric chloride contain four times 35.5 grammes of chlorine and 4 grammes of hydrogen, and as half its chlorine only is set free when it is acted upon by manganic binoxide, 146 grammes of hydric chloride only yield twice 35.5, or 71, grammes of free chlorine; so that to find how much hydric chloride must be taken to yield a given weight of free chlorine, say 50 grammes, the following statement must be worked out:—

Chlorine.		Grms. of hydric chloride.		Chlorine.
71	:	146	: :	50

When chlorine gas is passed into a solution of ammonia, it combines with the hydrogen of the ammonia, forming hydric chloride, setting the nitrogen free. This is a method of obtaining nitrogen, and is one which should be given when you are asked how nitrogen is prepared. It was not, however, mentioned when treating of nitrogen, as this seemed the better place for its consideration. While the ammonia is in excess

nitrogen is evolved. If the solution be concentrated, each bubble of chlorine produces a flash of light. When the chlorine is in excess, a substance called nitric chloride is formed: it is an oily liquid, heavier than water and of an extremely explosive character, owing to the loose way in which the nitrogen and chlorine are combined. It is said that the action of chlorine on ammonic chloride is to unite with the hydrogen, the nitrogen of the ammonic chloride being left in combination with chlorine. The exact constitution of this body is not well known, and it is of too dangerous a nature to admit of very accurate analysis. When heated to between 93°C. and 100°C. it explodes violently, breaking into pieces the vessel in which it is contained; even a slight concussion will cause its explosion.

Chlorine is able to decompose steam at a high temperature. If chlorine gas and steam be caused to pass through a tube heated to redness, hydric chloride is formed by the hydrogen of the steam uniting with it, and oxygen is set free.

Questions on Chlorine.

1. Describe the physical properties of chlorine. In what condition is it always found in nature? From what article of common use is it usually prepared?

2. What is the density of chlorine? What is its density with respect to atmospheric air?—*Ans.* 2.47.

3. Would it be advisable to collect chlorine over mercury or cold water? Give the reasons for your answer. What is the best method of collecting it?

4. At what pressure can chlorine be liquefied?

5. Give the easiest and simplest method of forming antimonic and cupric chlorides. Describe the phenomena attending their formation.

6. When a candle burns in chlorine, does chlorine combine with its carbon? Does it combine with its hydrogen?

7. What is the action of chlorine on moistened litmus paper, and what on a solution of indigo?

8. What would be the effect of chlorine gas on a bad-smelling drain?

9. How is chlorine most conveniently prepared? Describe fully the process of its preparation.

10. What weight of chlorine can be obtained from 1,752 grammes of hydric chloride? How much manganic binoxide must be used to liberate it?—*Ans.* 852 grammes of chlorine; 1,044 grammes of manganic binoxide.

11. When 29.2 grammes of hydric chloride are decomposed by manganic binoxide, how much chlorine is evolved, and how much manganic chloride is obtained?—*Ans.* 14.2 grammes of chlorine; 25.2 grammes of manganic chloride.

12. What volume will 124.25 grammes of chlorine occupy?—*Ans.* 39.2 litres.

13. How much manganic binoxide will 29.2 grammes of hydric chloride require for its decomposition?—*Ans.* 17.4 grammes of manganic binoxide.

14. 8.76 grammes of hydric chloride are to be decomposed; how much manganic binoxide must be taken? How much manganic chloride, water and chlorine will be formed?—*Ans.* 5.22 grammes manganic

binoxide, 7.56 grammes of manganic chloride, 2.16 grammes of water, and 4.26 grammes of chlorine.

15. Give a method for the preparation of nitrogen in which chlorine is employed.

16. What is the action of chlorine on steam at a high temperature?

Compounds of Oxygen and Chlorine.

These are five in number:—

- Hypochlorous acid, which with water forms hydric hypochlorite;
- Chlorous acid, which with water forms hydric chlorite;
- Chloric oxide, which forms no compound with hydrogen or other metals;
- Hydric chlorate, of which the acid has not been isolated;
- Hydric perchlorate, of which the acid has not been isolated.

HYPOCHLOROUS ACID.

Hypochlorous acid consists of chlorine and oxygen in the following proportions:—87 grammes of the acid occupy twice 11.2 litres, and contain twice 35.5 grammes of chlorine and 16 grammes of oxygen; its density therefore is half 87, or 43.5.

It is a yellow gas, and is obtained by passing chlorine gas over mercuric oxide. Mercuric oxide contains mercury and oxygen in the proportion of 200 grammes of

mercury to 16 grammes of oxygen, and when chlorine is made to act upon it, twice 35.5 grammes of chlorine unite with the 16 of oxygen, forming 87 grammes of hypochlorous acid, and twice 35.5 grammes more chlorine unite with 200 grammes of mercury to form 271 grammes of mercuric chloride, the chlorine taking the place of the oxygen of the mercuric oxide. Hypochlorous acid is easily condensed to a liquid by a mixture of ice and salt; in this condition its colour is red, and its oxygen and chlorine not being very firmly held together it readily explodes.

Hydric hypochlorite contains chlorine, oxygen and hydrogen. 35.5 grammes of chlorine, 16 grammes of oxygen, and 1 gramme of hydrogen give 52.5 grammes of hydric hypochlorite. It has but a slight acid reaction; carbonic acid, which is one of the weakest acids, is able to decompose it, and set hypochlorous acid free from any hypochlorite.

When chlorine gas is passed into water, a decomposition of part of the water takes place. Chlorine unites with half its hydrogen to form hydric chloride; the other half remaining with its oxygen also combines with chlorine to form hydric hypochlorite. If argentic nitrate be added to this mixture of hydric chloride and hydric hypochlorite, the silver of the argentic nitrate will unite, with all the chlorine of the hydric chloride, forming white insoluble argentic chloride, which will fall to the bottom of the vessel, while hydric hypochlorite will remain in solution along with hydric nitrate, formed by the hydrogen of the hydric chloride and the nitrogen and oxygen set free from the argentic nitrate; thus only half the chlorine passed into the water is in a state to

precipitate the silver. This circumstance should be remembered, as it will help you to understand a class of phenomena of considerable interest.

When 1 part of mercuric oxide is mixed with 12 parts of water, and submitted to the action of chlorine gas, mercuric oxychloride and hydric hypochlorite are formed, the mercuric oxychloride is precipitated as a brown powder, which contains mercuric oxide, united with mercuric chloride and hydric hypochlorite remains in solution. The hypochlorite has strong bleaching powers, and its concentrated solution is a powerful oxidising agent; as its chlorine and oxygen are not firmly held together, it easily gives up its oxygen to any body which is capable of taking it.

Bleaching powder, commonly called chloride of lime, is a combination of calcic oxide and hypochlorous acid, and therefore contains calcium, oxygen, and chlorine; 143 grammes of it contain 40 grammes of calcium, twice 16, or 32, grammes of oxygen, and twice 35.5, or 71, grammes of chlorine; half the oxygen, 16 grammes, originally belonged to the calcic oxide, and the other half to the hypochlorous acid, which, as we have seen, contains oxygen and chlorine in the proportion of 16 to 71.

Bleaching powder is made by passing chlorine gas over moist calcic oxide (lime); this is done in large chambers made of lead or Yorkshire flagstones. The lime is placed on perforated shelves, and the chlorine is let in slowly, otherwise too great a heat would be caused by the combination, and this would cause the formation of another body, called calcic chlorate, which would be useless for bleaching purposes. The action

which takes place is as follows: The lime being moistened is in the form of calcic hydrate, and contains calcium, oxygen and hydrogen; the chlorine decomposes the calcic hydrate, forming calcic chloride, calcic hypochlorite and water; half the chlorine taken up unites with calcium to form calcic chloride, the other half unites with half the oxygen, and the remaining half of the calcium of the calcic hydrate, to form calcic hypochlorite, and all the hydrogen combines with the remaining oxygen to form water.

Grms. of calcic oxide.	Grms. of chlorine.	Grms. of calcic chloride.	Grms. of calcic hypochlorite.	Grms. of water.
148	and 4 (35.5)	give 111	143	and 36

In commercial bleaching powder there is never the theoretical quantity of chlorine. Dilute acid liquids easily decompose bleaching powder, setting free hydric hypochlorite, which can be distilled over. If the acid be too strong, free chlorine is evolved. For bleaching purposes, a solution of bleaching powder is made, and the goods to be bleached are dipped into it, and are afterwards dipped into a solution of very dilute hydric sulphate; it is only when they have been dipped in the acid solution that the bleaching action takes place. This operation is called "souring," and it is by it that the bleaching powder in solution is decomposed, and hydric hypochlorite is set free in the fibres of the cloth, where it gives up its oxygen, which, by forming new compounds with the colouring matters, alters their nature, and makes them colourless, that is to say, hydric hypochlorite bleaches by oxidising substances, and its effects are permanent, as the colours bleached cannot be restored. The peculiar smell of bleaching powder is

well known in sick rooms. As hypochlorous acid is so weak an acid, it is liberated from bleaching powder by the carbonic acid in the air, in the presence of moisture, and it readily oxidises those substances which produce bad and unwholesome odours, forming with them innocuous compounds, and in this way it acts as a disinfecting agent.

Potassic chlorate.—We shall consider potassic chlorate next, as it is the source from which the other oxides of chlorine are obtained. When chlorine gas is passed into water, hydric chloride and hydric hypochlorite are formed. Water is hydric oxide, *i.e.* a compound of oxygen and hydrogen; if the hydrogen in water be replaced by potassium (as it can easily be), we have potassic hydrate, which, in many respects, behaves like water. Therefore we may expect that if chlorine gas is made to act on potassic hydrate, we shall get compounds similar to those which result from its action on water, and this is the case. When chlorine is passed into a solution of potassic hydrate, potassic chloride and potassic hypochlorite are formed, but this only takes place when the solution is *cold.* If the solution of potassic hydrate be *hot,* no potassic hypochlorite will be formed, but potassic chlorate; and if the *cold* solution, through which chlorine has been passed, be boiled, the potassic hypochlorite will be converted into potassic chlorate, for potassic hypochlorite is decomposed by heat into potassic chloride and potassic chlorate: the chlorate, as the termination *ate* implies, contains more oxygen than the hypochlorite. If three times 90.5 grammes of potassic hypochlorite be heated, twice 74.5 grammes of potassic chloride and 122.5 grammes of

potassic chlorate will be formed. If we compare the composition of potassic hypochlorite with that of potassic chlorate, the relative proportions of oxygen will be seen.

Grms. of potassic hypochlorite.		Grms. of potassium.	Grms. of chlorine.		Grms. of oxygen.
90.5	contain	39	35.5	and	16
Grms. of potassic chlorate.		**Grms. of potassium.**	**Grms. of chlorine.**		**Grms. of oxygen.**
122.5	contain	39	35.5	and	3 times 16, or 48

Potassic chlorate is not nearly so soluble in water as potassic chloride, so that after chlorine has been passed into the solution of potassic hydrate, and if (after the solution has been evaporated sufficiently to allow the potassic chlorate to crystallise), it be left to cool, the potassic chlorate crystallises out, leaving the potassic chloride in solution; and in this way potassic chlorate is made. If the crystals of the chlorate be washed and carefully dried and dissolved in distilled water and recrystallised, they will be free from any adhering chloride.

Hydric chlorate can be obtained from potassic chlorate by precipitating the potassium with hydric fluosilicate; the potassium takes the place of the hydrogen of the hydric fluosilicate, forming insoluble potassic fluosilicate, and the hydrogen replaces the potassium of the chlorate, forming hydric chlorate. Or, if hydric sulphate be added to a solution of baric chlorate an insoluble baric sulphate is precipitated, leaving hydric chlorate in solution; it is, however, dilute, but may be evaporated by a gentle heat, and a strong solution of hydric chlorate obtained. Hydric chlorate is a

very powerful oxidising agent; when dropped on paper, the oxidation of the paper is so violent that it bursts into flame. Various chlorates are formed by the replacement of the hydrogen of hydric chlorate by metals; all these are soluble, therefore hydric chlorate gives no precipitate by which its presence in a solution can be determined. In this, as in other respects, it resembles hydric nitrate.

Test for a Chlorate.

The way to test for a chlorate is to free the liquid first from all traces of chlorides; this is done by adding argentic nitrate till no further precipitate is formed; the precipitate is then to be filtered off, and into the liquid which passes through the filter should be put some metallic zinc and hydric sulphate; hydrogen will be evolved, and this will take the oxygen from the chlorate, and when all the oxygen is removed only a chloride will be left.

If any excess of argentic nitrate was added when the chlorides were first precipitated, it will, with the chloride formed from the chlorate, give a white precipitate of argentic chloride, or, if there be no free argentic nitrate present, the white precipitate may be obtained on the addition of it.

It is always well to add hydric nitrate to the solution after the action of the zinc and hydric sulphate, because if more hydrogen is evolved than is necessary to reduce the chlorate, it will reduce the argentic chloride, and the white precipitate formed, which is argentic chloride, will disappear; its chlorine unites with the zinc, me-

tallic silver being precipitated, and the hydric nitrate is added to dissolve this silver, which, when dissolved, again forms argentic chloride by taking the chlorine from the zinc chloride, and giving up its nitrogen and oxygen to the zinc to form zinc nitrate.

Chlorous acid is a yellowish green gas, easily decomposed, and is formed by the reduction or deoxidisation of hydric chlorate. For this purpose potassic chlorate is taken, and the agent employed for its deoxidisation is arsenious acid in the presence of hydric nitrate; the arsenious acid first takes oxygen from the hydric nitrate (forming hydric nitrite) and this takes oxygen from the chlorate (forming hydric nitrate and chlorous acid); this gas must be collected by downward displacement. When united with water it forms *hydric chlorite*, and different metallic chlorites can be formed by substituting other metals for hydrogen.

Chlorites are very easily decomposed by the weakest acids, even by carbonic acid. Hydric chlorite bleaches very readily. As its name implies, it contains less oxygen than hydric chlorate, and more than hydric hypochlorite. In 106.5 grammes of potassic chlorite there are 39 grammes of potassium, 35.5 of chlorine, and 32 of oxygen. Potassic chlorate contains 48 grammes of oxygen to 39 grammes of potassium, and the hypochlorite 16 grammes of oxygen to 39 of potassium. Both chloric and chlorous acids form only one class of salts, being in this respect like nitric acid.

Chloric oxide is a deep yellowish-coloured strong-smelling gas; it is obtained by gently heating potassic chlorate with hydric sulphate; it is extremely

explosive, and therefore its preparation is very dangerous. It can be condensed to a red liquid at a temperature of (— 20°C.) It has a strong bleaching action, and is very soluble in water. 135 grammes of chloric oxide contain 2(35.5) grammes of chlorine and 4(16) grammes of oxygen.

Hydric perchlorate.

Perchloric acid has not been obtained. "When potassic chlorate is heated, it first fuses and begins to give off oxygen; at a certain point, however, the whole mass solidifies; if the decomposition be stopped at this stage a new salt will be found to be contained in the residue, together with chloride and unaltered chlorate." It may be easily separated from the chlorate by the action of "hydric chloride which decomposes this latter, but has no action on the perchlorate."—*Roscoe.*

Potassic perchlorate is not readily soluble, and is therefore separated from the chloride by crystallisation in the same way as described for the separation of potassic chlorate from the chloride. Hydric perchlorate can be obtained from potassic perchlorate by distilling it with hydric sulphate.

It is a powerful oxidising agent, and immediately chars vegetable substances when brought in contact with them: 138.5 grammes of potassic perchlorate contain 39 grammes of potassium, 35.5 grammes of chlorine, and 64 grammes of oxygen; it therefore contains 16 grammes more oxygen in proportion to potassium than does potassic chlorate. There exists but one class of perchlorates just as there is but one class of chlorates and chlorites.

Questions on Hypochlorous Acid and Hydric Hypochlorite.

HYPOCHLOROUS ACID.

1. How can hypochlorous acid be obtained? What is its constitution?

2. What volume do 26.1 grammes of hypochlorous acid occupy?—*Ans.* 6.72 litres.

3. What is the colour of liquid hypochlorous acid, and what kind of union exists between its constituent elements?

HYDRIC HYPOCHLORITE.

1. What are the bodies formed when chlorine gas is passed into water? How would you prove that the chlorine in these bodies exists in different states?

2. In what way must mercuric oxide be used when it is employed in the manufacture of hydric hypochlorite?

3. In what consists the bleaching powers of hydric hypochlorite? Explain the bleaching action of common chloride of lime.

4. How much chlorine ought there to be in a kilogramme of chloride of lime, theoretically?—*Ans.* 496.5 grammes.

5. How much chlorine must be passed over 34.2 grammes of calcic hydrate to form 42.9 grammes of calcic hypochlorite?—*Ans.* 42.6 grammes.

6. If chlorine gas be passed over 171 grammes of calcic hydrate, how much calcic chloride and how much calcic hypochlorite ought to be formed if all the hydrate be acted upon?—*Ans.* 128.25 grammes of calcic chloride, and 165.223 grammes of calcic hypochlorite.

7. What is the difference between the action of chlorine on a hot and a cold solution of potassic hydrate?

8. How can potassic chlorate in solution be separated from potassic chloride?

9. In 42.875 grammes of potassic chlorate, how much chlorine is there, and how much potassium?—*Ans.* 13.65 grammes of potassium, and 12.425 grammes of chlorine.

10. How much potassic chlorate must be decomposed to yield 2.73 grammes of potassium, and 2.485 grammes of chlorine?—*Ans.* 8.575 grammes.

11. 4.344 grammes of potassic hypochlorite are boiled in water; how much potassic chloride and potassic chlorate will be formed?—*Ans.* 2.388 grammes of potassic chloride, and 1.967 grammes of potassic chlorate.

12. What are the two methods of making hydric chlorate, and what effects are produced by a strong solution of it?

13. Give instances of the resemblance between hydric chlorate and hydric nitrate in the constitution of their compounds.

14. Explain how to obtain a chloride from a solution of a chlorate. How is this made use of in testing for a chlorate?

15. How is chlorous acid obtained from potassic chlorate?

16. What is contained in the flask in which potassic chlorate has been heated, if the action has been stopped at the point when the chlorate solidifies?

17. How can potassic perchlorate be separated from potassic chlorate?

CHAPTER XIX

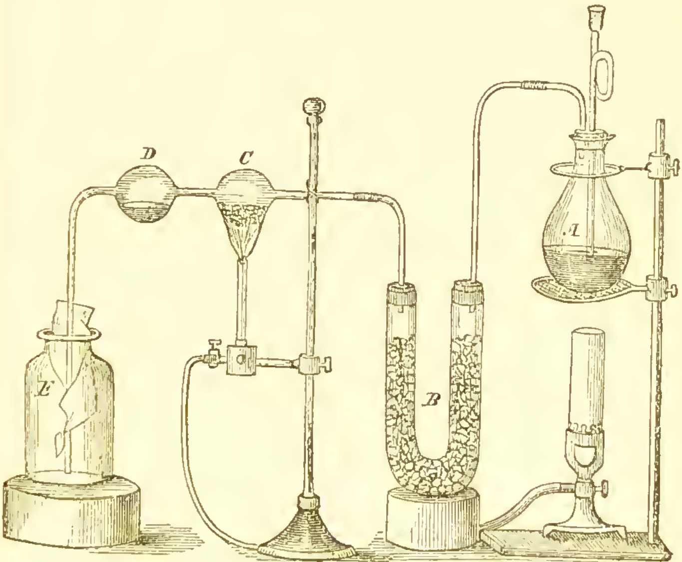
COMPOUNDS OF CHLORINE AND HYDROGEN—HYDRIC CHLORIDE.

Hydric chloride.—When 11.2 litres of chlorine are mixed with 11.2 litres of hydrogen, no change takes place if they are kept in the dark; but when exposed to daylight a union gradually takes place between the two gases, and twice 11.2 litres of hydric chloride are formed. In this case the volume is the same before and after union, *i.e.* there is no condensation. If the mixed gases be exposed to direct sunshine their union takes place with explosive violence. Two points here should be especially remembered, first that hydrogen and chlorine combine *directly*, and, secondly, *that their combination is not attended with any reduction in the volume* in which they are mixed; so that whatever be the volume of hydric chloride, one half that volume is formed by hydrogen and the other half by chlorine.

In 2 volumes of hydric chloride there is 1 volume of hydrogen and 1 volume of chlorine. That this is the case may be proved analytically, by collecting a volume of hydric chloride over mercury, and shaking it up with sodium amalgam; the chlorine unites with the sodium, forming sodic chloride, and the mercury rises in the glass tube to half the height at which it originally stood, the

gas now occupying the space above the mercury is hydrogen. Sodium amalgam is a compound of mercury with metallic sodium—the mercury of the sodium amalgam is set free, and forms no compound.—*Hofman.*

That chlorine and hydrogen exist in hydric chloride is well demonstrated by an experiment devised by Professor Graham. If hydric chloride be generated



in a vessel A, and be then passed through a tube B bent in the shape of the letter U and filled with fused calcic chloride in order to dry the gas; and if it then be passed through another tube in which two bulbs have been blown, in one of which, C, is placed some manganic binoxide, when this bulb is heated by the lamp beneath it, the hydric chloride passing over the hot binoxide

will be decomposed—manganous chloride will be formed by the union of its chlorine with the manganese, and water will be formed by the union of the oxygen of the manganic binoxide and with the hydrogen of the hydric chloride which will be collected in bulb D; but as only half the chlorine combines with the manganese, and the remainder is set free, this passing on into the bottle E will bleach the piece of moistened litmus paper placed within it.

11.2 litres of chlorine weigh 35.5 grammes, and as 11.2 litres of hydrogen weigh 1 gramme, therefore twice 11.2 litres of hydric chloride weigh 36.5 grammes, and so its density with respect to hydrogen is 18.25. Hydric chloride is a gas without colour; it is very soluble in water, of which 1 volume dissolves 454 volumes of the gas at (15°C.)—*Roscoe.*

In the form of a solution it is kept for use, and it is this solution which is usually called hydrochloric acid, and which was formerly called muriatic acid, and spirits of salt. Hydric chloride has been condensed to a liquid at a pressure of 40 atmospheres.

Hydric chloride has strongly acid properties; it reddens blue litmus paper, and when mixed with alkaline bodies, which turn red litmus paper blue, it destroys their alkaline properties and loses its own acid properties by forming with them compounds which have no action on either blue or red litmus, and they are therefore called neutral.

When hydric chloride is acted on by most of the metals, they unite with the chlorine, and hydrogen is evolved. Thus, zinc and hydric chloride form zinc chloride and free hydrogen, just as zinc and hydric

sulphate formed zinc sulphate and free hydrogen. We see, therefore, that the sulphur and oxygen, in hydric sulphate, behave in the same way as does the chlorine in the hydric chloride.

When a metallic oxide is acted upon by hydric chloride a chloride of the metal is formed, and the oxygen of the oxide unites with the hydrogen of the hydric chloride to form hydric oxide (water). The usual way of making hydric chloride is by acting upon sodic chloride (common salt) with hydric sulphate; an interchange takes place similar to that which occurs in the manufacture of hydric nitrate—the sodium of the sodic chloride takes the place of the hydrogen in the hydric sulphate, forming sodic sulphate, and this hydrogen combines with the chlorine of the sodic chloride to form hydric chloride.

As in the manufacture of hydric nitrate, the operation consists of two stages, one half the hydrogen only is displaced at first, hydro-sodic sulphate being formed, and this at a very high temperature reacts on more sodic chloride and the second half of the hydrogen is replaced by sodium, more hydric chloride being made. To prevent frothing in the retort, and to get rid of moisture, the common salt is fused in a crucible; this renders it more dense; it is then poured out on a stone slab to cool, after which it is broken in pieces, and mixed with the hydric sulphate in a retort, and heat is applied gently. The hydric chloride comes over and is collected like ammonia gas in a bottle containing distilled water which absorbs it readily, forming a fuming solution which contains about 43 per cent. of hydric chloride. Previous to collection it is passed

through a bottle containing a small quantity of water, in order to arrest any impurities which might spirt over from the hot liquid in the retort.

An apparatus, similar to that used in the preparation of ammonia may be employed in making it on a small scale. To this apparatus, however, a safety funnel, such as is used in the last diagram, must be added, as dangerous explosions happen when water is sucked back into hot hydric sulphate. In large quantities it is made in iron vessels, where the heat can be raised sufficiently high to effect the complete displacement of hydrogen by sodium in the hydric sulphate; this cannot be done in a glass retort, therefore the solid residue which remains behind when it is made in such a vessel is hydro-sodic sulphate. Large quantities of hydric chloride in an impure state are made in the manufacture of sodic carbonate from sodic chloride. The impurities of this acid are iron, which causes it to have a yellow colour, free chlorine, sometimes hydric sulphite and hydric sulphate, and almost invariably arsenic and sodic chloride.

Tests for Chlorides.

If argentic nitrate be added to a soluble chloride, a white precipitate of argentic chloride is formed, which is insoluble in hydric nitrate, but soluble in ammonia, from which solution in ammonia it is reprecipitated by hydric nitrate.

If manganic binoxide be added to a chloride, together with hydric sulphate, and if heat be applied, chlorine

gas will be evolved, which may be distinguished by its odour, and by its action in bleaching moistened litmus paper.

For the application of the following test, if the chloride be in solution it must be evaporated to dryness, strong hydric sulphate must be added, and some crystals of a substance called potassic dichromate; this experiment should be performed in a test tube. On the application of heat a deep red gas will be evolved; this gas is called chlorochromic acid, and when conducted into water is decomposed; a pale yellow solution being formed, from which the colour is not discharged by ammonia. This permanence of the yellow colour serves to distinguish between chlorine and bromine as will be further explained in treating of bromine.

Questions on Hydric Chloride.

1. Describe the method of forming hydric chloride from chlorine and hydrogen. In what proportions must the gases be mixed?

2. If 51 litres of hydrogen be mixed with 51 litres of chlorine, and be caused to combine, what will be the volume of hydric chloride formed?—*Ans.* 102 litres.

3. How many litres of hydrogen are there in 20 litres of hydric chloride?—*Ans.* 10 litres.

4. How can you prove that there is as much hydrogen by volume, as chlorine, in hydric chloride?

5. Describe Professor Graham's experiment proving the composition of hydric chloride.

6. What are the results of the action of zinc on a solution of hydric chloride?

7. Describe the usual methods of preparing hydric chloride. What is the residue if the operation be conducted in a glass retort? And what if it be carried on in a vessel heated to a high temperature?

8. What are the impurities of common hydric chloride, and from what source do they emanate?

9. What weight of hydric chloride can be obtained from 409.5 grammes of sodic chloride, if it be completely decomposed by hydric sulphate, and what volume will it occupy?—*Ans.* 255.5 grammes, and 156. litres.

10. What volume do 177.5 grammes of chlorine occupy? And what weight of hydrogen must be added so as to produce a mixture in the proportions for making hydric chloride?—*Ans.* 56 litres of chlorine and 5 grammes of hydrogen.

11. What weight of sodic chloride must be decomposed in order to produce 365 grammes of hydric chloride, and what weight of hydric sulphate must be used to effect complete decomposition?—*Ans.* 585 grammes of sodic chloride, and 980 grammes of hydric sulphate.

12. It is required to decompose 234 grammes of sodic chloride, what weight of hydric sulphate must be taken? What will be the weight of the salt remaining in the retort? And what the volume of the gas liberated?—*Ans.* 392 grammes of hydric sulphate, 480 grammes of hydro-sodic sulphate, and 89.6 litres.

13. From what weight of sodic chloride can 50 litres of hydric chloride be obtained?—*Ans.* 130.58 grammes.

14. Argentic nitrate is added to a soluble chloride, what kind of precipitate is formed? What is its com-

position? State the action on it of hydric nitrate and of ammonia.

15. In what way can chlorine be obtained, as such, from a solution of a chloride? If present in small quantities how would you detect it?

16. When you act on the dry residue, obtained from evaporating a solution, with potassic dichromate and hydric sulphate, and you obtain a deep red-coloured gas, is that sufficient proof that a chloride was in the residue? If not, explain how you would confirm the presence of a chloride.

CHAPTER XX.

BROMINE.

Bromine, the second element in this group, receives its name from its offensive odour; it is derived from the word Βρῶμος, *bromos*, a stink. It is found in sea water, and was discovered in 1826 by Balard. After other salts, which crystallise more easily, have been separated from the sea water, the bromine remains behind, combined with sodium, magnesium, potassium, &c., forming compounds called bromides; they are, therefore, more soluble than sulphates and chlorides. From this liquid, which is commonly called the mother liquid of sea water, or bittern, because it has produced the crystals already formed, bromine is prepared in the following manner. Chlorine is passed through the mother liquid and the bromine is set free; for chlorine is chemically more powerful than bromine, and can free it from those compounds which are called bromides (it cannot, however, free it from other compounds called bromates, which contain oxygen, and which will be described presently).

Most bromides are white, and when dissolved in water are colourless, the chlorine changes the liquid to a brown colour, and this is owing to the pre-

sence of free bromine, which is brown, or some say yellowish brown. The solution of bromine is then shaken up with ether, which dissolves the bromine, and after the bottle containing it has stood quiet for some time, the ether, with the bromine dissolved in it, rises and floats on the surface, for ether is much lighter than water, the ether solution of bromine is then drawn off from the top, or the aqueous solution is allowed to flow out from a tap in the bottom of the vessel. The ether, containing bromine, is then shaken up with a solution of sodic or potassic hydrate. As the bromine now unites with the sodium or potassium to form a bromide and bromate, the liquid becomes colourless. The bromide of sodium or potassium is then mixed with manganic binoxide and hydric sulphate (the ether having been first evaporated by heat and collected in a condenser for future use), and heat is gradually applied; a decomposition similar to that produced by the action of manganic binoxide and hydric sulphate on sodic chloride takes place, and bromine is given off; it is condensed in a cooled receiver.

If a bromide, say sodic bromide, be heated with strong hydric sulphate some bromine will be set free; this is not the case when sodic chloride is heated with hydric sulphate, for then hydric chloride is obtained, as we have already seen, and none of this is decomposed by the strong hydric sulphate, but hydric bromide, which is formed under similar circumstances, is in part decomposed by strong hydric sulphate at a high temperature,—thus proving that chlorine is chemically more powerful than bromine.

At the ordinary temperature bromine is a liquid of a deep brown-red colour. (Chlorine is a greenish-yellow gas.) A bottle containing bromine has generally some of the vapour of bromine above the liquid, showing that it is easily volatilised or changed into vapour. It has a strong pungent odour, and irritates the air passages of the lungs when inhaled even in very small quantities. Minute traces of it mingled with air have very much the odour of decaying seaweed.

Bromine vapour is more than twice as heavy as chlorine. 11.2 litres of bromine vapour weigh 80 grammes, whereas the same measure of chlorine weighs 35.5 grammes. Its density with respect to hydrogen is therefore 80; its specific gravity (in the liquid form) that is, its weight compared with an equal bulk of water, is 2.731.—*Miller*. Bromine dissolves in water, but very slightly. It is dissolved readily by alcohol and by ether.

OXIDES OF BROMINE.

Bromine combines more powerfully with oxygen than does chlorine, and forms with it three compounds.

Hydric hypobromite.—Hypobromous acid corresponding to hypochlorous acid has not been obtained, but its hydrogen salt, hydric hypobromite, can be made by treating bromine water with mercuric oxide. Hydric hypobromite has a bleaching action similar to hydric hypochlorite, that is, it bleaches by oxidation; but as bromine holds oxygen in combination more powerfully than chlorine does, it parts with it less easily, and therefore it does not bleach so readily. In

97 grammes of hydric hypobromite there are 80 grammes of bromine, 1 of hydrogen, and 16 of oxygen.

Hydric bromate.—Bromic acid has not been obtained. Potassic bromate is made by adding bromine to potassic hydrate; and proceeding on the same method as described in the manufacture of potassic chlorate. Strong hydric sulphate when heated with potassic bromate decomposes it, bromine and oxygen being set free.

Hydric bromate may also be made by passing chlorine gas through bromine water, and as bromine unites more readily with oxygen than chlorine does, hydric chloride and hydric bromate are produced. 129 grammes of hydric bromate contain 80 grammes of bromine, 1 gramme of hydrogen, and 48 grammes of oxygen.

So little is known about **Hydric perbromate** that its consideration is hardly necessary for our present object.

Hydric bromide.—Bromine and hydrogen do not unite as readily as chlorine and hydrogen, that is, they do not unite in diffused daylight nor under the influence of direct sunshine. If, however, hydrogen and bromine vapour be passed through a tube heated to redness hydric bromide is formed. This behaviour of bromine should be remembered, as it shows that its hydrogen compound is less firm and stable than that of chlorine with hydrogen.

Hydric bromide is a gas, and if 2 volumes of hydric bromide be decomposed, they will yield 1 volume of bromine vapour and 1 volume of hydrogen; that is, in 22.4 litres of hydric bromide there are 11.2 litres

of bromine vapour and 11.2 litres of hydrogen; and as 11.2 litres of bromine vapour weigh 80 grammes, and 11.2 of hydrogen weigh 1 gramme, 22.4 litres of hydric bromide will weigh 81 grammes and its density is 40.5 compared with that of hydrogen.

Hydric bromide can also be made by decomposing phosphoric bromide with a small quantity of water. Phosphoric bromide and water form hydric bromide and hydric phosphate. The hydrogen of the water in part unites with the bromine of the phosphoric bromide, and its remaining hydrogen and all its oxygen form, with the phosphorus, hydric phosphate.

Hydric bromide has strongly acid properties, and resembles hydric chloride in its behaviour with metals and metallic oxides.

Tests for Bromides.

A solution of argentic nitrate gives a precipitate of argentic bromide when added to a soluble bromide. Argentic bromide is not as soluble in ammonia as is argentic chloride, so that if a solution of argentic chloride in ammonia be added to a soluble bromide, a precipitate of argentic bromide will be formed.

If a soluble bromide be treated with chlorine water, bromine is set free, and can be separated out with ether in the manner already described.

If bromine be mixed with hydric sulphate and potassic dichromate no compound is formed by the bromine and chromic acid, analogous to chlorochromic acid, but red fumes pass over when heat is applied. These dissolve in water, forming a yellow liquid, and

might lead to the supposition that chlorine was present; but the addition of ammonia decolorises the bromine solution, for it is simply the bromine dissolved in water which colours it, and ammonia unites with the bromine, forming a bromide, which in solution is colourless, whereas in the reaction with a chloride the yellow solution formed is one of chromic acid, which forms a yellow compound with ammonium.

Questions on Bromine.

1. From what sources is bromine obtained? How can free bromine be obtained from sodic bromide?

2. What is the action of hydric sulphate on potassic bromide? Compare it with the action of hydric sulphate on sodic or potassic chloride.

3. What effect has chlorine on solutions of bromides?

4. If bromine water be shaken up with ether what effect will be produced on the colour of the liquid?

5. What effect would potassic hydrate produce on a solution of bromine? State both the chemical and physical effects.

6. What is the density of bromine vapour?

7. Which combines most powerfully with oxygen, bromine, or chlorine? Give proofs in support of your answer.

8. Explain why hydric hypobromite does not bleach as readily as hydric hypochlorite.

9. What quantity of oxygen is there in 20.37 grammes of hydric hypobromite?—*Ans.* 3.36 grammes.

10 How much oxygen and hydrogen are necessary to

unite with 16.8 grammes of bromine to form hydric hypobromite?—*Ans.* .21 grammes of hydrogen and 3.36 of oxygen.

11. What quantity of oxygen, hydrogen, and bromine can be obtained from 4.074 grammes of hydric hypobromite? — *Ans.* 3.36 grammes of bromine, .042 grammes of hydrogen, and .672 grammes of oxygen.

12. Describe the action of chlorine gas on bromine water.

13. Under what conditions do bromine and hydrogen unite directly?

14. Phosphoric bromide is acted upon by water, state the decomposition which occurs.

15. State the characters in which hydric bromide resembles hydric chloride, and in what it differs from it.

16. What volume do 58.32 grammes of hydric bromide occupy?—*Ans.* 16.128 litres.

17. How many volumes of bromine vapour are there in 6 volumes of hydric bromide? And what is their weight with respect to the weight of the hydrogen contained in them?—*Ans.* 3 volumes of bromine vapour; 80 times greater.

18. How would you detect a bromide? Describe the difference between argentic chloride and argentic bromide.

CHAPTER XXI.

IODINE.

Chlorine is a gas, bromine a liquid, but iodine is a solid at the ordinary temperature. It is much heavier than bromine; 11.2 litres of its vapour weigh 127 grammes; its density, therefore, compared with hydrogen, is 127.

Iodine is found in the ashes of seaweeds, this substance, from which it is obtained, is called kelp. It is also found in some minerals, and has lately been discovered by Dr. Charles Graham to exist, in no inconsiderable quantities, in phosphorite, in Rhenish Prussia.

In 1811 it was accidentally found by Courtois in the manufacture of sodic carbonate from kelp. It is now extracted by treating the ashes of seaweeds in the following manner. After the ashes have been heated gently for some time with water, and their solution concentrated, it is allowed to cool, and the more insoluble soda and potash salts crystallise out. The iodine remains in the mother liquid, which is treated with a small quantity of hydric sulphate, and which, after standing for some hours, to allow a further crystallisation to take place. is strained off, and placed in

a retort with manganic binoxide, and more hydric sulphate. The application of heat causes a reaction to take place similar to that in the manufacture of chlorine, and iodine is set free, which is condensed in a series of glass receivers. Iodine is a crystalline metallic-looking solid of a dark blackish colour; but its vapour is of a most beautiful violet colour, hence its name, from the Greek word *ἰώδης*, *iodes*, violet.

Iodine has an odour which bespeaks its marine origin, being somewhat similar to that of both chlorine and bromine, though not so pungent; it is slightly soluble in water, with which it forms a brown solution. It is much more readily soluble in alcohol, ether and a solution of potassic iodide, and to these solvents it also gives a brown colour; but when dissolved in carbonic sulphide the solution is violet, similar to the colour of iodine vapour.

Iodine is immediately set free from its compounds called iodides, which do not contain oxygen, by bromine and chlorine; and in its free state, if starch be added to it, it forms a blue solution if the iodine be present in even the smallest quantities; but when present in larger quantities the colour formed is almost black.

The specific gravity of iodine is 4.95; it melts at 107°C. and boils at 175°C. If chlorine be used to set iodine free from iodides care must be taken not to have the chlorine in excess, as it destroys the blue colour formed by iodine and starch, the best reagent for this purpose, when testing for iodine, is nitrous acid gas. It should be generated from hydric nitrate and copper in a separate test-tube; the gas, being heavier than air, admits of its being poured like water. If

a small quantity be poured into the tube containing the supposed iodide, and be shaken up with the solution, the iodine will be immediately set free, and can be distinguished either by the solution turning brown, or if present in small quantities, by the blue colour produced on the addition of starch.

OXIDES OF IODINE.

Iodic acid, which also forms a hydrogen compound called hydric iodate. Hydric periodate is also known. There is no compound analogous to hypochlorous acid.

Iodic acid is obtained by boiling iodine with hydric nitrate. It crystallises readily from a concentrated aqueous solution—*Williamson.* In 334 grammes of iodic acid there are 2(127), or 254 grammes of iodine, and 5(16) or 80 grammes of oxygen. Heat decomposes iodic acid, iodine and oxygen being set free.

Hydric iodate.—If solid iodine be put into water, and if chlorine gas be passed into the mixture, the iodine disappears and hydric iodate is formed, together with hydric chloride. Iodine combining more powerfully with oxygen than does chlorine, the water gives up one sixth of its hydrogen and all its oxygen to the iodine, and the remaining five sixths of its hydrogen unites with chlorine to form hydric chloride; thus—

Iodine 127 grammes, water 3(18) grammes, chlorine 5(35.5) grammes, give hydric iodate 176 grammes, hydric chloride 182.5 grammes.

Potassic iodate is made by adding iodine to

a solution of potassic hydrate, potassic iodide and potassic iodate are formed; or, better, when chlorine is passed into a mixture of iodine and potassic hydrate, potassic iodate and potassic chloride are formed.

As iodine unites more readily than chlorine with oxygen, if a solution of potassic chlorate be treated with iodine the iodine takes the place of the chlorine, forming potassic iodate. It is in this way that **potassic periodate** is made, for it is formed when iodine acts on potassic perchlorate.

Questions on Iodine.

1. How does iodine occur in nature? Describe its extraction from kelp.

2. What are the appearances of iodine as a solid and iodine as a vapour?

3. What is the difference between the solution of iodine in alcohol and iodine in carbonic sulphide?

4. When chlorine gas is passed through water in which iodine is suspended, what bodies are formed?

5. How much iodine is there in $58.\dot{6}$ grammes of hydric iodate?—*Ans.* $42.\dot{3}$.

6. What is the difference between the action of chlorine and of iodine on a cold solution of potassic hydrate? Explain the reason of this difference.

7. What is the action of iodine on a solution of potassic chlorate?

8. How much free iodine can be obtained from .498 grammes of potassic iodide?—*Ans.* .381 grammes.

COMPOUND OF HYDROGEN AND IODINE.

Hydric iodide.—Iodine, like chlorine, combines with hydrogen; but its compounds with hydrogen are less stable than those of chlorine, or bromine with hydrogen. Iodine does not unite directly with hydrogen. It should be remembered that chlorine unites directly with hydrogen even in diffused daylight; that bromine unites with it, neither in daylight nor in direct sunlight, but only when heated with it to a high temperature; and that iodine does not unite with it directly under any circumstances.

Hydric iodide is a colourless gas, soluble in water; 128 grammes of it occupy 2(11.2) litres; its density therefore is 64, that of hydrogen being 1. 1 gramme of hydrogen, occupying 11.2 litres, is united with 127 grammes of iodine, occupying 11.2 litres, in hydric iodide.

There are two ways in which hydric iodide can be easily prepared: one is by passing hydric sulphide (called sulphuretted hydrogen gas) through iodine suspended in water; sulphur is precipitated, and the hydrogen of the hydric sulphide unites with the iodine to form hydric iodide; thus—

Iodine 2(127) grammes and hydric sulphide 34 grammes, give sulphur 32 grammes, and hydric iodide 2(128) grammes.

The second method is by acting on water with phosphoric teriodide. In this case hydric iodide and hydric phosphite are formed. Phosphoric teriodide contains

31 parts of phosphorus to 3(127) parts of iodine, therefore—

Phosphoric teriodide 412 grammes, and water 54 grammes, give hydric iodide 3(128) grammes, and hydric phosphite 82 grammes.

Potassic iodide.—When iodine in excess is boiled with a solution of potassic hydrate, potassic iodide, potassic iodate and water are formed. If the mixture be gently evaporated to dryness, excess of iodine being present, the residue will consist of nothing but potassic iodide and potassic iodate. If this be heated strongly, the oxygen is driven off from the iodate, and nothing remains but potassic iodide.

It should be remembered here that chlorine and bromine when mixed with potassic hydrate form, in the cold, respectively, potassic chloride, and potassic hypochlorite, or potassic bromide, and potassic hypobromite, and that, after boiling, the potassic hypochlorite or hypobromite is changed, the one into potassic chloride and chlorate, the other into potassic bromide and bromate; but with iodine no such change takes place, for we have seen that iodine forms no compound with oxygen similar to a hypochlorite or hypobromite.

The impurities to which potassic iodide is liable are free potassic hydrate (which is present if the iodine was not kept in excess during the evaporation of the solution containing potassic iodide and iodate), potassic iodate, and free iodine, for iodine is as before stated very soluble in potassic iodide. Potassic carbonate is also sometimes present. If on the addition of hydric acetate effervescence takes place, it proves

the presence of a carbonate. If the liquid which, when pure, is colourless, turns brown, it proves the presence of iodate, for when potassic iodide and iodate are together, a weak acid substance like hydric acetate will set iodine free; though, when separate, iodine is liberated from neither of them by strong hydric nitrate or hydric chloride if these acid liquids be pure.

Tests for Iodides.

The presence of free iodine manifests itself at once by the brown colour which it gives to the solution. The starch test for iodine has been already explained; there is yet another and very useful one which should be remembered.

Argentic nitrate gives with a soluble iodide a yellowish-white precipitate of argentic iodide, which is insoluble in hydric nitrate, also in ammonia, but which turns white on addition of ammonia. Argentic chloride, you will remember, is readily soluble in ammonia; argentic bromide is soluble, but not so soluble as argentic chloride; and argentic iodide is absolutely insoluble in ammonia.

Questions on Hydric and Potassic Iodide.

1. In what respects does iodine differ from chlorine and bromine in its behaviour towards hydrogen?

2. Does iodine burn in oxygen?

3. Describe the physical properties of hydric iodide. How can it be made? What is the decomposition which occurs when hydric sulphide is employed? And what when phosphorus iodide?

4. Hydric sulphide is passed through water holding in suspension .762 grammes of iodine, how much hydric iodide will be formed, and how much sulphur will be deposited?—*Ans.* .768 grammes of hydric iodide and .096 grammes of sulphur.

5. 12.36 grammes of phosphoric teriodide are decomposed by water, how much hydric iodide and how much hydric phosphate will be formed?—*Ans.* 11.52 grammes of hydric iodide and .249 grammes of hydric phosphate.

6. How much water must be taken to decompose 20.6 grammes of phosphoric teriodide, and what weight of hydric iodide will be formed?—*Ans.* 2.7 grammes of water and 19.2 grammes of hydric iodide.

7. How many litres of hydric sulphide are required to form hydric iodide when passed into water containing 50.8 grammes of iodine in suspension?—*Ans.* 4.48 litres.

8. It is desired to make 6.272 litres of hydric iodide, what volume of hydric sulphide will be required, and what weight of iodine?—*Ans.* 3.136 litres of hydric sulphide and 35.56 grammes of iodine.

9. How much iodine and how much potassium is there in 5.03 grammes of potassic iodide? — *Ans.* 1.181 grammes of potassium and 3.848 grammes of iodine.

10. What is the effect of a red heat on a mixture of potassic iodide and potassic iodate?

11. Explain why there is no iodine compound with oxygen and potassium similar to potassic hypochlorite.

12. To what impurities is potassic iodide liable, and how do they come to be present?

13. If potassic iodide contain potassic iodate, how can the presence of the iodate be detected?

14. Describe the method of making potassic iodide, and what precaution should be taken to prevent the presence of free potassic hydrate?

15. A solution contains an iodide, how would you distinguish the precipitate given on the addition to it of argentic nitrate, from that which argentic nitrate gives with a soluble chloride?

16. Write down in the order of their solubility in ammonia, the silver salts of bromine, iodine and chlorine.

CHAPTER XXII.

FLUORINE.

Fluorine is found in nature combined with calcium, as calcic fluoride, or fluor spar, commonly called Derbyshire spar. It is not known in the free state, although it is said to have been obtained free by the action of iodine on argentic fluoride in a sealed glass tube. Fluorine forms no compound with oxygen, but with hydrogen it forms a colourless gas called hydric fluoride.

Hydric fluoride is made by acting on fluor spar (calcic fluoride) with hydric sulphate. It is found that if 78 grammes of calcic fluoride be mixed with 98 grammes of hydric sulphate, and heated gently, 40 grammes of hydric fluoride and 136 grammes of calcic sulphate are formed; the hydrogen of the hydric sulphate unites with the fluorine, and the calcium unites with the sulphur and oxygen of the hydric sulphate.

It will be seen that there were in the calcic fluoride 40 grammes of calcium, for when the calcium changed places with the 2 grammes of hydrogen in hydric sulphate, the calcic sulphate was found to weigh 136 grammes, that is, 38 grammes more than it did before the interchange of hydrogen for calcium took place;

and as the calcic fluoride weighed 78 grammes, therefore there were in it 38 grammes of fluorine which, by the interchange, took into combination with them 2 grammes of hydrogen, and therefore 19 parts, by weight, of fluorine unite with 1 part of hydrogen.

It is found that 20 grammes of hydric fluoride occupy 2(11.2) litres, and that the hydrogen in it occupies 11.2 litres; therefore we conclude that 19 grammes of fluorine occupy 11.2 litres, and that the density of fluorine is 19.

Hydric fluoride dissolves in water, forming a corrosive, fuming, very acid liquid; the hydrogen in it can be replaced by metals forming fluorides, such as potassic fluoride, argentic fluoride, &c., the hydrogen being set free. When metallic oxides are dissolved in it—the same fluorides are formed, but the oxygen of the oxide unites with the hydrogen of the hydric fluoride forming water, so that in this case no free hydrogen is given off.

Hydric fluoride must be made in platinum or lead vessels—because it attacks silica (or flint) which is an invariable constituent of glass, and therefore it destroys glass vessels. From its possessing this peculiar property, it is used for etching on glass. The patterns on glass in public-house windows are etched out with this liquid.

If a piece of glass be coated with wax, and a design be scratched out in the wax leaving the glass exposed, and if a solution of hydric fluoride be placed on the glass, after a short time the pattern will be found to be eaten out, for the wax is not affected by hydric fluoride.

Glass thermometers are marked in this way, but the vapour is generally used for the purpose and not the solution. Hydric fluoride has been condensed to a liquid at a temperature of —20°C. but it is doubtful whether dry hydric fluoride has been obtained in the liquid state.—*Roscoe.*

Fluorine forms a gaseous compound with silicon, which will be explained when the element silicon is treated of. When hydric fluoride falls on the skin it causes very unpleasant wounds, unless it be very quickly washed off. As it destroys glass it is generally kept in gutta percha bottles (which it does not destroy), and sometimes in lead vessels.

Tests for a Fluoride.

The presence of a fluoride may be detected by placing some of the substance suspected to contain it in a small platinum vessel or common lead inkstand, along with some hydric sulphate, and then putting over it a piece of glass coated with wax, through which figures have been traced; the application of a very gentle heat will cause hydric fluoride to be formed, if a fluoride be present, and it will manifest its presence by etching the figures traced in the wax on the surface of the glass.

QUESTIONS ON FLUORINE.

1. Describe the peculiar characteristics of fluorine, stating the differences between it and iodine.

2. How does fluorine occur in nature?

3. Give the composition of hydric fluoride, and state the method of its preparation.

4. How much fluorine is needed to combine with .2 grammes of hydrogen to form hydric fluoride?—*Ans.* 3.8 grammes.

5. How much hydric fluoride is evolved, and how much calcic sulphate is formed, when 15.6 grammes of calcic fluoride are heated with 19.6 grammes of hydric sulphate?—*Ans.* 8 grammes of hydric fluoride and 27.2 grammes of calcic sulphate.

6. How much calcic fluoride must be acted upon by hydric sulphate to form 1,280 grammes of hydric fluoride?—*Ans.* 2,496 grammes.

7. What volume do 4 grammes of hydric fluoride occupy at the standard temperature and pressure?—*Ans.* 4.48 litres.

8. Why is it necessary to use platinum or lead vessels in making hydric fluoride?

9. Describe the method of testing for a fluoride.

CHAPTER XXIII.

SULPHUR.

Sulphur is found native, that is, uncombined with other substances, crystallised in yellow crystals in a form which is called octahedral; it occurs in this state in volcanic districts. It also exists largely in combination with metals forming sulphides—as in iron pyrites, copper pyrites, zinc blende, galena, &c. Zinc blende is zinc sulphide; and galena is plumbic or lead sulphide. It also exists in nature in the form of sulphates, which, as you know, contain oxygen as well as sulphur. Sulphate of barium, or barytic sulphate, is found in large quantities; it is called heavy spar: also calcic sulphate, or gypsum, from which plaster-of-paris is made.

Sulphur is an important constituent of animal bodies; it is found in albumen. The white of egg is albumen, and it is the presence of sulphur in it which causes silver spoons to get black when they are used as egg spoons.

Sulphur is largely obtained from iron pyrites, from which it is driven off by heat, the air being excluded, otherwise it would unite with the oxygen of the air and form a gaseous compound with it

Impure sulphur, after importation into this country, is refined by heating it in closed vessels made of earthenware; it still, however, retains volatile impurities like arsenic—which is pretty constantly present in sulphur.

When sulphur is heated in a closed vessel and its vapour passes into another vessel which is kept cold, it is condensed in the form of small bright yellow particles; in this form it is sold in commerce under the name of flowers of sulphur.

Sulphur is also used in medicine in the form of precipitated sulphur. This is a yellowish white powder, very fine, and of a dirty appearance. It is made by precipitating sulphur from some of its compounds, as when hydric sulphide is passed into chlorine water. The most common form of sulphur is that called roll sulphur, which is made by melting sulphur and casting it in moulds. Sticks of rolled sulphur crackle when gently heated; if held in the hand, its warmth will cause the sulphur to crackle. When sulphur is heated it melts at 115°C.; if heated beyond this point, it forms a thick brown treacle-like liquid, and it continues to get thicker, so that the vessel containing it may be inverted without its running out. The addition of more heat causes it to become fluid again; this happens at about 250°C., and at about 485°C., or 490°C., it boils, and is converted into a red vapour.

If sulphur melted in a common iron ladle be allowed to cool, and if the crust which forms on the top of the mass be broken and that portion which is liquid be poured out, the cavity will be found to have beautiful needle-shaped crystals of sulphur projecting into it. These

crystals are called prismatic; they have a different specific gravity from the octohedral crystals before mentioned. The specific gravity of the octohedral crystals is 2.05, and that of the prismatic 1.98. In a short time, the prismatic crystals become opaque, and break up into small octohedra. If sulphur be heated to about 230°C., and be then poured into water, it forms a mass which is somewhat elastic and which can be drawn into threads, and has an appearance similar to gutta percha softened in hot water; after a short time, however, it again assumes its ordinary brittle condition.

Sulphur does not dissolve in water; it dissolves in carbonic sulphide (bisulphide of carbon), in chloride of sulphur, and in benzole; also in hot turpentine.

When a solution of sulphur in carbonic sulphide is allowed to evaporate, crystals of octohedral sulphur are left. Sulphur, while in the tenaceous state, obtained by pouring it, when melted, into water, is not soluble in carbonic sulphide. These various forms in which sulphur exists are termed allotropic modifications. They are usually said to be three in number: first, the octohedral, or natural variety; second, the prismatic; and third, the viscid, or tenaceous.

Sulphur resembles oxygen in many respects. It is by chemists classed in a group of which oxygen is the first member. Carbonic acid contains carbon and oxygen; carbonic sulphide, also called sulphocarbonic acid, contains sulphur in the place of the oxygen; water, hydric oxide, contains hydrogen and oxygen; hydric sulphide, which contains hydrogen and sulphur, may be regarded as water in which the oxygen is replaced by sulphur.

Sulphur combines directly with most metals. If sulphur be vapourised, copper, and some other metals, take fire in it and burn, as metals burn in oxygen.*

Questions on Sulphur.

1. How is sulphur found in nature? In what animal substances does it occur?
2. From what source is sulphur generally obtained? and state the method used for its extraction.
3. What is the difference between flowers of sulphur and precipitated sulphur? State the method of preparing both kinds.
4. How can prismatic crystals of sulphur be obtained? Is this form permanent?
5. Describe the different "allotropic" modifications in which sulphur exists.
6. Compare oxygen and sulphur.

Compounds of Oxygen with Sulphur.

Sulphur forms very stable compounds with oxygen; more stable than those of oxygen with chlorine, bromine, and iodine which we have been considering. These compounds are two in number, but there are seven compounds of sulphur, oxygen and hydrogen. Two only will need notice at present.

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* The vapour density of sulphur will be given in Part II.

Sulphurous acid.—When sulphur is burnt in air or oxygen gas it burns with a blue flame, forming sulphurous acid, which is a gas at the ordinary temperature and pressure, colourless, and having a smell which is easily recognised as the odour of burning brimstone. This gas is liquefied by a pressure of 2 atmospheres, or at a temperature of —10°C., and it becomes solid at —76°C. 32 grammes of sulphur combine with 2(16) grammes of oxygen to form 64 grammes of sulphurous acid, and this weight of the gas measures 2(11.2) litres; its density, therefore, with respect to hydrogen is 32.

Sulphurous acid can be prepared by heating hydric sulphate with mercury; the sulphate is decomposed, mercuric sulphate, water and sulphurous acid being formed. Charcoal also, by the aid of heat, decomposes hydric sulphate, but the sulphurous acid formed in this way is mixed with carbonic acid. Sulphurous acid is readily soluble in water, and this solution of it is used in the laboratory and in manufactories. When sulphurous acid is dissolved in water, the solution may be regarded as one of hydric sulphite. You will remember that the termination *ite* is used for compounds formed by acids whose names end in *ous*, and which contain less oxygen than those acids whose names end in *ic*.

Sulphurous, like carbonic acid forms two classes of compounds, for 82 grammes of hydric sulphite contain 2 grammes of hydrogen, 32 grammes of sulphur, and 3(16) grammes of oxygen; and these 2 grammes of hydrogen can be replaced separately by potassium, so

that hydropotassic sulphite, which contains 1 gramme of hydrogen and 39 grammes of potassium, can be formed; also potassic sulphite, which contains twice 39 grammes of potassium and no hydrogen. Hydropotassic sulphite, in which only half the hydrogen of hydric sulphite has been replaced by potassium, has acid properties, whereas potassic sulphite is neutral.

Sulphites are decomposed by bodies with stronger acid properties, such as hydric nitrate, hydric sulphate, and hydric chloride, with evolution of sulphurous acid. Sulphites on exposure to air become slowly oxidised into sulphates.

Sulphurous acid is used for bleaching purposes; its action in such cases is called a reducing action, for it is said to take away oxygen from the colouring substance; its action, therefore, is exactly opposite as a bleaching agent to that of chlorine, which, as you will remember, bleaches by oxidation. Sulphurous acid is principally used for bleaching woollen and silk goods, plaited straw goods, and others which might be injured by chlorine.

Articles which have been bleached with sulphurous acid may have their colour restored by the use of alkalies; whereas a colour destroyed by chlorine can never be restored. Sulphurous acid has powerful antiseptic properties; it prevents putrefaction, and is used in the preservation of animal substances.

Test for Sulphurous Acid.

The presence of sulphurous acid is detected in the following manner. If barytic chloride be added to a

solution containing a sulphite, a white precipitate is formed, soluble in hydric chloride, and if to this solution of barytic sulphite in hydric chloride, chlorine or bromine water be added, a white precipitate of barytic sulphate is thrown down, which is insoluble in hydric chloride or hydric nitrate. Here the chlorine or bromine water has oxidised the sulphite into a sulphate by the process of oxidation already described in the chapter on chlorine.

Questions on Sulphurous Acid.

1. How is sulphurous acid prepared for laboratory uses?

2. If carbon is used for the preparation of sulphurous acid, to what impurity is it liable? and how can that impurity be got rid of?

3. How many compounds are there of potash and sulphurous acid?

4. How much oxygen is there in 3347.2 grammes of sulphurous acid?—*Ans.* 1673.6.

5. What volume do 19.2 grammes of sulphurous acid occupy?—*Ans.* 6.72 litres.

6. Describe the method of detecting a sulphite in a solution; and explain the action of the reagents employed.

SULPHURIC ACID.

Sulphuric acid, also called sulphuric trioxide, as its name implies, contains more oxygen than sulphurous acid. 80 grammes of sulphuric acid contain 32 grammes of sulphur, and 3(16) grammes of oxygen,

and these 80 grammes, in the state of vapour, occupy 2(11.2) litres; its density, therefore, with respect to hydrogen is 40. If sulphur be burnt in oxygen at a very high temperature, sulphuric, and not sulphurous, acid is formed, for at high temperatures sulphurous acid in the presence of oxygen is oxidised into sulphuric.—*Williamson.*

Sulphuric acid is usually prepared by heating Nordhausen acid in a retort, and condensing the white fumes which come over in a receiver cooled with ice and salt. These fumes form white silky crystals which deposit on the sides of the receiver, and when in mass they have the appearance of asbestos; this substance may be handled if the fingers be dry.

When thrown into water it unites with it with such energy that a hissing noise is produced, and great heat evolved. Sulphuric acid can also be formed by acting upon hydric sulphate with phosphoric acid; the phosphoric acid takes away from it the elements of water. Thus, hydrogen 2 grammes, sulphur 32 grammes, oxygen 4(16) grammes, form 98 grammes of hydric sulphate; and if from this 2 grammes of hydrogen and 16 grammes of oxygen (which are the proportions of oxygen and hydrogen in water) be taken away, 80 grammes of sulphuric acid, *i.e.* 32 grammes of sulphur and 3(16) grammes of oxygen, will remain.

Sulphuric acid cannot be obtained from hydric sulphate by the action of the most intense heat. If 2 volumes of sulphuric acid be passed through a tube heated to a high temperature, they are decomposed into sulphurous acid and oxygen—2 measures or volumes of sulphurous acid, and 1 of oxygen.

It may here be well to state the reason why the term acid is applied to bodies which have been for some time, and still are by some, called anhydrous acids, or anhydrides. It is admitted by leading chemists that what are usually called acids are hydrogen salts, compounds formed with water by, what are termed, anhydrides, *i.e.* acids. When the hydrogen of these salts is in part replaced, as we have seen it can be, by metals, such as potassium, which is chemically more powerful than the metal hydrogen, the acid properties are in part neutralised, but when all the hydrogen is replaced by potassium the acid is entirely neutralised, and we shall see, when we come to the consideration of phosphates, that when all the hydrogen of trihydric phosphate is replaced by potassium the salt is alkaline; we therefore conclude that the body that has the power of conferring these acid properties in various degrees, according to the strength of the metallic oxide with which it combines, is really the acid although it cannot be placed in a position in which it can manifest acid properties; for sulphuric acid, cannot redden blue litmus paper, because it, being a solid, cannot come into contact with it; and if water be used to dissolve the acid, it immediately becomes hydric sulphate. Tartaric acid, *as it is called*, if dry, does not redden litmus paper; it requires to be dissolved that contact may take place. The great difficulty here seems to lie in the two uses which are made of the word acid. To pursue this subject further at present would be beyond our purpose; it will be resumed when the composition of salts is considered.

Questions on Sulphuric Acid.

1. What is the composition of sulphuric acid? and describe its properties.
2. How is sulphuric acid best obtained?

HYDRIC SULPHATE.

Hydric sulphate, called also sulphuric acid and oil of vitriol, is a thick oily liquid, colourless when pure, but frequently brownish, from the presence of organic matter. 98 grammes of it contain 2 grammes of hydrogen, 32 grammes of sulphur, and 4(16) of oxygen. Experiments have been made which show that, at the temperature at which its vapour volume has been taken, hydric sulphate separates into sulphuric acid and water, and that this causes the volume to be double that which the vapour of other compounds occupies. It appears, therefore, that at present the vapour volume of hydric sulphate cannot be taken.

Hydric sulphate has a powerfully corrosive action; it destroys organic substances. It takes up water (with avidity), and is therefore used in chemical operations, particularly in drying gases; it takes away from many compounds, which contain the elements of water, their oxygen and hydrogen in the proportion in which they form water. Thus, if hydric sulphate be poured into a syrup made with sugar, it takes away the oxygen and hydrogen which in sugar exist in the proportions to form water, and leaves behind the carbon of the sugar as a black residue. Again, with the aid of

heat hydric oxalate is decomposed by hydric sulphate, all its hydrogen and part of its oxygen in the proportions in which they exist in water are retained by it, while its carbon and the remainder of its oxygen unite to form carbonic acid and carbonic oxide, which are set free. Hydric sulphate is made by the oxidation of sulphurous acid and the addition of water.

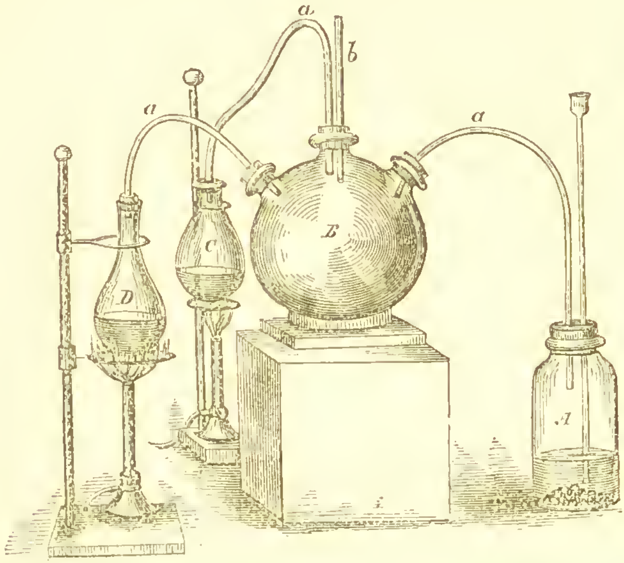
You will remember that when nitric oxide comes in contact with free oxygen it unites with it, forming nitrous acid, which is a red gas; this again easily gives up this additional oxygen to bodies which are capable of taking it. Sulphurous acid has this power, and when it meets with nitrous acid takes from it a portion of its oxygen, and becomes sulphuric acid, the nitrous acid becoming nitric oxide again, and the sulphuric acid meeting with water becomes hydric sulphate. Thus, 64 grammes of sulphurous acid, which contain 32 grammes of sulphur and 2(16) grammes of oxygen, when they meet with 76 grammes of nitrous acid, which contain 2(14) grammes of nitrogen and 3(16) grammes of oxygen, take from them 16 grammes of oxygen, forming 80 grammes of sulphuric acid, and this, with 18 grammes of water, forms 98 grammes of hydric sulphate.

Hydric sulphate is used largely in the arts, and is manufactured in the following manner: Sulphur or iron pyrites is burnt in a furnace through which air passes freely, the oxygen of the air unites with the sulphur, forming sulphurous acid. In another small furnace sodic nitrate is heated with some oil of vitriol. Hydric nitrate is thus set free, and the sulphurous acid, coming in contact with the hydric nitrate at a high temperature, takes from it part of its oxygen,

and nitric oxide is formed. These combined gases pass into large leaden chambers, where the nitric oxide takes oxygen from the air and becomes nitrous acid; it gives this oxygen to the sulphurous acid, thus converting it into sulphuric acid, itself becoming nitric oxide, and this nitric oxide repeats the operation with more air and sulphurous acid which enter the chamber from the furnace where the sulphur or iron pyrites is being heated; thus nitric oxide acts as a carrier of oxygen, first taking it from the air, and then giving it up to the sulphurous acid; jets of steam are continually forced into the chambers under pressure from a boiler beneath them. In the bottom of the leaden chambers where the gases mix, water is placed for the purpose of condensing the hydric sulphate as it is formed. The steam unites with the sulphuric acid, produced by the oxidation of the sulphurous acid, forming hydric sulphate, which is condensed and dissolved by the water in the bottom of the chambers. The solution of hydric sulphate thus formed is dilute, though sufficiently strong for some purposes. When required of greater strength, it is evaporated in open leaden vessels until it reaches the specific gravity 1.720. It can be still further concentrated by heating it in glass or platinum retorts, and in this way it can be obtained of specific gravity 1.842, its boiling point being 327°C.

Hydric sulphate can be prepared in laboratories in the following manner:—A glass globe, B, having three necks or openings, is placed on a stand. In a bottle, A, nitric oxide is evolved by the decomposition of hydric nitrate with copper, and the gas is led into the globe B by the glass tube *a*. D is a flask connected by a tube *a*

with the receiver B; in this, sulphurous acid is made by either of the methods already given. C contains water, from which steam is obtained, and conducted by the tube *a* into the vessel B. *b* is an outlet tube.



Hydric sulphate forms two crystalline compounds, with water, containing different proportions of oxygen and hydrogen, which will be further noticed in Part II.

Commercial oil of vitriol, *i.e.* impure hydric sulphate, contains several impurities; lead in small quantities, which it dissolves from the leaden vessels in which it is evaporated; this impurity may be detected by diluting it with water, for then plumbic sulphate, which is more soluble in strong than in dilute hydric sulphate, is precipitated as a white powder. Arsenic, which you will remember is an impurity of sulphur, and which is volatile at a tolerably high temperature, passes

into the chambers where the hydric sulphate is made, mixes with it, and forms one of its most constant impurities. The method of detecting this impurity is too complicated to be described here. Hydric nitrate and perhaps some of the other oxides of nitrogen are frequently found in hydric sulphate; they may be detected by the tests already described for nitrates. Sulphurous acid and even hydric chloride are also at times present.

Sulphuric acid, like carbonic acid, forms two classes of compounds, for the two parts of hydrogen which its hydrogen salt contains may be separately replaced by other metals, such as sodium or potassium; as we have seen in the preparation of hydric nitrate we can have hydro-potassic sulphate and potassic sulphate.

Tests for a Sulphate.

Barytic chloride gives, with hydric sulphate, or any soluble sulphate, a white precipitate of barytic sulphate, which is insoluble in all acid liquids.

Questions on Hydric Sulphate.

1. When sulphuric acid unites with water, what body is formed? What are its properties?

2. For what purposes is hydric sulphate employed for which its readiness to combine with water peculiarly fits it?

3. Describe the method of manufacturing oil of vitriol, and explain fully the office performed by nitric oxide.

4. What are the common impurities of hydric sulphate? Which impurity is discovered by simple dilution?

5. How many grammes of oxygen must be added to 40 grammes of sulphurous acid to form sulphuric acid? and how much water to convert the sulphuric acid into hydric sulphate?—*Ans.* 10 grammes of oxygen and 11.25 grammes of water.

6. What is the weight of each of the constituent elements in 29.4 grammes of hydric sulphate?—*Ans.* hydrogen .66 grammes, oxygen 19.2 grammes, sulphur 9.6 grammes.

7. 102.4 grammes of sulphur are oxidised completely in the presence of moisture, how much hydric sulphate will they form?—*Ans.* 313.6 grammes.

8. How much sulphurous acid is required to make 3,283 grammes of hydric sulphate?—*Ans.* 2,144 grammes.

9. What effect does a soluble salt of barium produce when added to a solution containing a sulphate?

10. How many kinds of sulphates does potassium form when acted on by hydric sulphate? and state how they may be obtained.

COMPOUNDS OF SULPHUR AND HYDROGEN.

Hydric sulphide, otherwise called sulphuretted hydrogen, contains sulphur and hydrogen in the following proportions:—in 34 grammes of hydric sulphide

there are 32 grammes of sulphur and 2 grammes of hydrogen, and this weight occupies 2(11.2) litres, its density therefore when compared with hydrogen is 17.

Hydric sulphide is a colourless transparent gas, with an odour resembling that of rotten eggs; it burns with a pale greyish white flame. The products of its combustion in air are water and sulphurous acid; it can be condensed to a liquid at a pressure of about 16 atmospheres; it is soluble in water, which, at zero Centigrade, dissolves 4.37 times its volume of the gas.

Hydric sulphide is found free in volcanic districts, and in solution in waters which are frequently used as medicines. The Harrogate waters are of this kind, and their offensive smell shows the presence of hydric sulphide. It is also evolved by the decompositions which take place in drains and cesspools. If breathed, it is very unwholesome, even in small quantities, but in large quantities it is poisonous. Hydric sulphide is prepared by acting upon ferrous sulphide with dilute hydric sulphate, ferrous sulphate being formed and hydric sulphide being set free.

Another way of preparing it is by heating black antimonious sulphide with hydric chloride; in this case antimonious chloride is formed and hydric sulphide is set free. In evolving hydrogen gas, if strong hydric sulphate be poured upon the zinc before water is added, some sulphurous acid is formed, and this, when the water is added and hydrogen is evolved, is decomposed by the hydrogen, with formation of water and hydric sulphide; for sulphurous acid contains sul-

phur and oxygen, and these elements when respectively combined with proper proportions of hydrogen form those two compounds. Hydric sulphide is decomposed by chlorine, hydric chloride being formed and sulphur being set free.

Hydric sulphide is used in the laboratory in precipitating metals, from solutions of their salts, as sulphides, and from their behaviour with it the metals can be classed in different groups. The first group consists of those metals whose sulphides can be precipitated in the presence of free hydric chloride, and it contains the metals arsenic, antimony and tin, silver, mercury, lead, gold, platinum, bismuth, copper, and cadmium. The second group of metals, which consists of iron, manganese, nickel, cobalt, uranium, and zinc, are precipitated as sulphides in alkaline solutions; aluminium and chromium, which belong to this group are precipitated by hydric sulphide, not as sulphides, but as hydrated oxides. If hydric sulphide be passed into solutions containing barium, magnesium, calcium, strontium, potassium, and sodium, these metals are not precipitated.

Tests for Hydric Sulphide.

The tests for hydric sulphide are as follow: Add to the liquid containing the sulphide some hydric chloride; hydric sulphide will be set free, and may be recognised by its odour; if, however, the quantity be too small to be smelt, a piece of filter paper moistened with a solution of plumbic acetate should be held over the mouth of the test tube, and though a very small quantity of

hydric sulphide be evolved, it will blacken the paper. Hydric sulphide in very minute quantities gives a beautiful purple colour, with sodic nitro-prusside n presence of excess of ammonia.

Questions on Hydric Sulphide.

1. Describe the physical properties of this body How does it occur in nature?

2. How is hydric sulphide usually prepared?

3. If hydric sulphide be passed into chlorine water, what changes would take place?

4. State into what groups metals are divided according to their action with hydric sulphide, *i.e.* as to the conditions under which they form precipitates.

5. What are the products of the combustion of hydric sulphide?

6. How much oxygen do 67.2 litres of hydric sulphide require for their combustion? What weights of the products of their combustion will be formed?—*Ans.* 100.8 litres of oxygen, 54 grammes of water, and 192 grammes of sulphurous acid.

7. How much sulphur is required to make 1,054 grammes of hydric sulphide?—*Ans.* 992 grammes.

8. What volume do 1,054 grammes of hydric sulphide occupy?—*Ans.* 694.4 litres.

9. A sulphide in solution is given, how would you set free hydric sulphide? and how prove its presence?

COMPOUND OF SULPHUR AND CARBON.

Carbonic sulphide, or sulphocarbonic acid, also called bisulphide of carbon.—76 grammes of this substance contain 12 grammes of carbon and 2(32) grammes of sulphur, and occupy in the state of vapour 2(11.2) litres; its density, therefore, with respect to hydrogen is 38. If the vapour of sulphur be passed over red-hot charcoal, carbonic sulphide is formed. It is a heavy liquid, insoluble in water, with an extremely unpleasant odour; it burns with a bluish flame; it combines with the sulphides of many metals, forming compounds called sulphocarbonates; it dissolves gums, phosphorus, sulphur, and india-rubber.

Sulphur forms two compounds with chlorine.

CHAPTER XXIV.

PHOSPHORUS.

Phosphorus exists in two well-defined forms. The first is a waxy-looking substance, of a pale yellowish-white colour; it is usually seen in the form of round sticks, which are kept under water, as when exposed to air it readily gives off white fumes, which result from its combination with atmospheric oxygen. In the dark these fumes are luminous.

If phosphorus be heated to a temperature of 240°C. in an atmosphere of hydrogen or carbonic acid, that is to say, not in the presence of free oxygen, it is gradually changed into a red powder, which is called *red* or *amorphous phosphorus*. This variety is not chemically so active as the other; it does not fume in air, nor unite with oxygen while in this condition. If a piece of clear phosphorus be placed in a solution of cupric sulphate (sulphate of copper), which is a greenish blue liquid, the colour will gradually disappear and metallic copper will be precipitated on the phosphorus. This is owing to the phosphorus combining with the oxygen, with which the copper was united, the copper being therefore set free, as metal. Red phosphorus is unable to produce this effect on cupric sulphate.

Phosphorus is employed in the manufacture of

lucifer matches; the clear variety was formerly used, but it was found to cause a peculiar disease in the jaw-bones of the workmen. The red variety is now coming into use, as it does not produce that disorder, and from its comparative inactivity is less dangerous.

If red phosphorus be heated above the temperature at which it is formed, in atmospheric air, it is converted into the other variety and burns.

Clear phosphorus is soluble in carbonic sulphide; therefore, if any yellow phosphorus remains in the manufacture of red phosphorus, it can be separated by this solvent, which does not dissolve red phosphorus. Clear phosphorus is also soluble in caustic soda, decomposing it, and forming with it a gas and hydric hypophosphite; it has no such action on red phosphorus.

Red phosphorus may also be prepared by dissolving ordinary clear phosphorus in carbonic sulphide, and mixing with it a small portion of iodine, and keeping it for some hours in a sealed tube at a temperature of 100°C.

The specific gravity of clear phosphorus is 1.83; that of red phosphorus 2.1. The specific heat of red phosphorus is 0.1700; that of clear phosphorus being 0.1887. These two forms in which phosphorus exists are termed allotropic.

Phosphorus is not found uncombined in nature; it exists, in the form of phosphates. Phosphate of lime is found in the bones of animals, of which it forms a very large proportion; it is also found in other parts of the animal body, as well as in plants. Phosphorite, which is a phosphate of lime, is found in considerable quantities in Northern Germany.

Bone-earth, which is obtained from bones by burning

them in air, consists mainly of calcic phosphate (phosphate of lime). To obtain phosphorus from bone-earth it is heated with hydric sulphate diluted with water; a part of the lime is precipitated as calcic sulphate, and the rest remains in solution as a double phosphate of hydrogen and calcium, which is commonly called superphosphate of lime. If this superphosphate be evaporated to dryness and mixed with charcoal and heated in a retort, carbonic oxide is formed, and the vapour of phosphorus distils over and is condensed in water.

Phosphorus is cast in sticks, the operation being performed under water. If 310 grammes of calcic phosphate be mixed with 196 grammes of hydric sulphate, 272 grammes of calcic sulphate will be formed, and 234 grammes of calcic superphosphate; and from this calcic superphosphate 62 grammes of phosphorus can be obtained.

Questions on Phosphorus.

1. Under what circumstances is clear phosphorus converted into the amorphous variety? And why does not this change take place when it is heated in air?

2. When clear phosphorus is placed in a solution of cupric sulphate, describe the decomposition which takes place. Will red or amorphous phosphorus produce the same effect?

3. How is phosphorus found in nature? Describe the method by which it is obtained in the free state. Why, in casting sticks of phosphorus, is the operation conducted under water?

4. How much phosphorus can be obtained from 62 grammes of calcic phosphate?—*Ans.* 12.4 grammes.

5. How much phosphorus is contained in 53.82 grammes of calcic superphosphate?—*Ans.* 14.26.

6. What quantity of hydric sulphate must be employed to get the phosphorus from 38.13 grammes of calcic phosphate? and how much phosphorus can be obtained from them?—*Ans.* 24.108 grammes of hydric sulphate, and 7.626 grammes of phosphorus.

COMPOUNDS OF PHOSPHORUS AND OXYGEN.

Phosphoric acid, or phosphoric pentoxide.—When phosphorus is burnt in excess of air or oxygen, phosphoric acid is formed; it is a white powder, volatile at a high temperature, and combines readily with water—it is therefore used in certain circumstances instead of hydric sulphate or calcic chloride as a drying or dessicating agent. It is also able to take away the elements of water from substances which contain them. You will remember that hydric sulphate was able to do this, as was instanced in the decomposition of sugar by that substance; phosphoric acid does the same thing, but even with greater energy. 142 grammes of phosphoric acid contain 2(31) grammes of phosphorus and 5(16) grammes of oxygen.

Questions on Phosphoric Acid.

1. What is the product of the combustion of phosphorus in excess of air or oxygen? State its uses, properties, and composition.

2. What quantity of phosphorus must be burnt to form 46.86 grammes of phosphoric acid? How much oxygen will it take into combination?—*Ans.* 20.46 grammes of phosphorus, and 26.40 grammes of oxygen.

Trihydric phosphate.—When phosphoric acid comes in contact with water, the combination is so energetic that it produces a hissing sound, and trihydric phosphate is formed; this has been usually called phosphoric acid.

Trihydric phosphate contains hydrogen, phosphorus, and oxygen, in the following proportions:—98 grammes of it contain 3 grammes of hydrogen, 31 grammes of phosphorus, and 4(16) grammes of oxygen.

For general use it is prepared in a pure state by boiling phosphorus in hydric nitrate diluted with four or five times its weight of distilled water in a glass retort. The phosphorus disappears, and is oxidised by the hydric nitrate, which is decomposed, giving off red fumes of nitrous acid and peroxide of nitrogen. When the whole of the phosphorus disappears the liquid should be transferred from the retort to a platinum dish, small quantities of hydric nitrate being added from time to time to ensure the perfect oxidation of the phosphorus. The evaporation should be continued until the liquid in the dish becomes viscid. This operation cannot be performed in porcelain dishes because the phosphate attacks the glaze.

Trihydric phosphate can also be prepared from bone-earth. After bone-earth has been treated with hydric sulphate, as in the preparation of phosphorus if

the solution of calcic superphosphate be filtered off from the calcic sulphate, and the clear liquid be treated with ammonia, the lime of the superphosphate will be precipitated as tricalcic phosphate, and triammonic phosphate will remain in solution; and if the solution be filtered from the insoluble tricalcic phosphate, and be evaporated and heated to redness, a substance called glacial phosphoric acid will remain, the ammonia being driven off by the heat; and if this glacial phosphoric acid be boiled for some time in water, trihydric phosphate will be formed.

This hydric phosphate has been termed orthophosphoric acid; it gives, with argentic nitrate in a neutral solution, a yellow precipitate of triargentic phosphate.

In both methods of manufacturing trihydric phosphate which have been given, after evaporation in the platinum vessel, the viscid substance which remains (commonly called glacial phosphoric acid) is of a different composition from trihydric phosphate; it contains less hydrogen and oxygen, for in 80 grammes of it there are 1 gramme of hydrogen, 31 grammes of phosphorus, and 3(16) grammes of oxygen.

If we refer back to the proportions of these elements in trihydric phosphate, we shall find that the sum of their weights in grammes is 98; and if from this we subtract the weight of the elements of water taken in grammes, that is, 18 grammes, we shall obtain 80 grammes, *i.e.* the weight of glacial phosphoric acid. This acid body should be termed metaphosphate; the reasons for this will be given in the second part of this book. It gives in neutral solutions with argentic nitrate a white precipitate; it also coagulates albumen, such as

white of egg, and in this, as well as in other respects, resembles hydric nitrate. By long boiling with water it takes up the elements of water, forming trihydric phosphate.

The 3 grammes of hydrogen in 98 grammes of trihydric phosphate can be separately replaced by 23 grammes of sodium, forming three sodic compounds, which are termed—

Dihydro sodic phosphate, in which 1 gramme of hydrogen is replaced by 23 of sodium. This substance has acid properties.

Hydro disodic phosphate, commonly called rhombic phosphate of soda, in which 2 grammes of hydrogen are replaced by 2(23) grammes of sodium. This compound is neutral.

Trisodic phosphate, in which 3 grammes of hydrogen are replaced by 3(23) grammes of sodium. This substance is alkaline. There is a salt formed by replacing in trihydric phosphate 1 gramme of hydrogen by 23 grammes of sodium, and another gramme of hydrogen by 18 grammes of ammonium, and this is called microcosmic salt. When trihydric phosphate is neutralised by ammonia, and a solution of magnesia is added, a white crystalline precipitate is formed, which is soluble in a solution of hydric chloride; and this precipitate consists of trihydric phosphate, in which ammonium and magnesium have replaced the hydrogen.

Tests for Phosphates.

In a solution containing a soluble triphosphate, argentic nitrate gives a yellow precipitate of triargentic

phosphate. As it is difficult to get a liquid perfectly neutral, the best way to perform this test, if the solution contain little of the phosphate, is to pour a weak solution of ammonia, after the addition of the argentic nitrate, gently on the surface of the liquid in the test tube, and, as in testing for nitrates with ferrous sulphate, a ring will be formed at the junction of the liquids. This ring will be yellow.

The ammonio magnesian test has also been already alluded to; the best way to perform it is as follows:—Add to a solution of magnesic sulphate some ammonia solution; a white precipitate will be formed; dissolve it in a small quantity of aqueous hydric chloride, and add ammonia in excess. Render the solution to be tested slightly ammoniacal, and pour into it the magnesian solution; a white crystalline precipitate will be formed. If the quantity of phosphate present be small, the precipitate does not appear at once; scratching the inside of the test tube with a glass rod hastens its formation.

Test with Ammonic Nitro-Molybdate.

In applying this test, the solution of the phosphate should be made slightly, but decidedly, acid with the hydric nitrate, and ammonic nitro-molybdate should be added in quantity nearly equal to that of the liquid to be tested. If much phosphate be present, a yellow precipitate appears immediately; if not, it can be obtained on boiling. This precipitate is soluble in ammonia.*

* *Tetrahydric diphosphate*, called also pyrophosphoric acid, will be considered in Part II.

Questions on Trihydric Phosphate.

1. If phosphorus be burnt in a bell jar of oxygen standing over water, what substance will be formed?

2. How can pure trihydric phosphate be obtained?

3. If trihydric phosphate be heated to a red heat for some time, what compound is left? State what different chemical properties it possesses.

4. How can the trihydric phosphate be obtained from a solution of calcic superphosphate?

5. How many grammes of hydrogen can be replaced by sodium or potassium in 1,470 grammes of trihydric phosphate?—*Ans.* 45 grammes.

6. If 29.4 grammes of trihydric phosphate be heated to redness till it ceases to lose weight, what will be the weight of the residue? What elements will it have lost, and what weight of each of them?—*Ans.* 24 grammes: .6 grammes of hydrogen, 4.8 of oxygen.

7. How much sodium must replace hydrogen in 1,470 grammes of trihydric phosphate to convert it into rhombic phosphate of soda?—*Ans.* 690 grammes.

8. How many different phosphates are there derived from trihydric phosphate which contain sodium? Describe their different properties.

9. Describe the tests for a phosphate, and state particularly which are applicable to acid, and which to alkaline solutions.

10. Describe the method of testing for phosphates.

PHOSPHOROUS ACID.

Phosphorous acid.—110 grammes of phospho-

rous acid contain 2(31) grammes of phosphorus and 3(16) grammes of oxygen. It is formed when phosphorus is burnt in a limited supply of air. It unites with water, forming hydric phosphite.

Hydric phosphite is better made by acting upon phosphorous chloride with water. In this decomposition hydric phosphite and hydric chloride are formed; the hydric chloride can be got rid of by boiling.

In 82 grammes of hydric phosphite there are 3 grammes of hydrogen, 31 grammes of phosphorus, and 3(16) grammes of oxygen. In this salt, which has acid properties, only 2 parts of the hydrogen can be replaced by other metals, so that sodic phosphite contains 1 part of hydrogen, 2(23) parts of sodium, 31 parts of phosphorus, and 3(16) parts of oxygen; whereas in trihydric phosphate all the 3 parts of hydrogen can be replaced by sodium or other metals.

Hydric phosphite acts as a reducing agent, and precipitates gold, silver, and mercury, as metals, from their solutions.

Questions on Phosphorous Acid.

1. If 12.4 grammes of phosphorus be burnt in a limited supply of air, what will be the weight of the residue, and what is it called?—*Ans.* 22 grammes.

2. How much oxygen is required to oxidise 46.5 grammes of phosphorus into phosphorous acid? What weight of acid will be formed? and how much oxygen is required to convert it into phosphoric acid?—*Ans.* 36 grammes of oxygen, 82.5 grammes of phosphorous acid, and 24 grammes of oxygen.

3. How much hydrogen can be replaced by potassium in 82 grammes of hydric phosphite?—*Ans.* 2 grammes.

4. What is meant by the term, a reducing agent?

COMPOUNDS OF PHOSPHORUS AND HYDROGEN.

Hydric phosphide, or phosphuretted hydrogen, is a gas containing phosphorus and hydrogen in the following proportions:—34 grammes of it contain 31 grammes of phosphorus and 3 of hydrogen, and this weight occupies 2(11.2) litres. When hydric phosphite is heated it decomposes, forming hydric phosphate and hydric phosphide; in this way then, hydric phosphide can be made. It is, however, usually prepared by heating phosphorus in a solution of sodic hydrate: the gas so made ignites on coming in contact with atmospheric air; this peculiarity seems to result from the presence of a small quantity of another hydrogen compound of phosphorus. The gas prepared by heating hydric phosphite does not take fire spontaneously.

QUESTIONS ON HYDRIC PHOSPHIDE.

1. What volume do 170 grammes of hydric phosphide occupy?—*Ans.* 112 litres.

2. How much phosphorus is there in 5.1 grammes of hydric phosphide?—*Ans.* 4.65 grammes.

3. Describe two methods of preparing hydric phosphide. In which method is the hydric phosphide spontaneously inflammable?

CHAPTER XXV.

BORON.

Boron is found in nature combined with oxygen. This compound is called boracic acid; 70 grammes of it contain 2(11) grammes of boron and 3(16) grammes of oxygen. It is also found as borax, which is a compound of boracic acid and soda. Boron is best obtained by heating boric chloride or boracic acid with sodium. It is a blackish powder, and if it be heated to a high temperature with the metal aluminium it is dissolved, and crystallises out on cooling. If the aluminium be dissolved with dilute hydric chloride the crystals of boron remain; they are tolerably transparent, but are coloured brown. In this operation copper coloured scales of graphitoid boron are formed; they are opaque and like graphitoid carbon; their form is that of six-sided plates.

BORACIC ACID.

Boracic acid is the only known compound of boron with oxygen. It is volatile, when heated in the presence of moisture. In parts of Tuscany it occurs native, and escapes from the earth with jets of steam; hydric sulphide is generally present with it, as might be expected, it being a volcanic product. These steam jets containing boracic acid are called soffioni or fume-

rolles; their products are collected in excavations in the mountain, and from them borax is made. The heat evolved by the fumerolles is used for evaporating the liquid from which the borax crystallises.

Boracic acid is obtained by heating a solution of borax with hydric sulphate. Hydric borate crystallises out from this solution. If to a solution of borax a small quantity of hydric sulphate be added, it will take soda from part of the borax, leaving hydric borate free. If blue litmus paper be dipped into this solution it will be changed by the hydric borate to a purple colour, indicating its slightly acid properties. Yellow turmeric paper, which is turned brown by alkaline liquids, when dipped into the same solutions and dried, is changed to a brown tint. Boron combines directly with nitrogen at a red heat, forming boric nitride. Chlorine combines with boron, forming boric chloride, which is easily decomposed by water.

Tests for Borates.

Boracic acid imparts a green colour to flame. The best test for its presence is the following:—Add to a solution, supposed to contain it, a small quantity of sodic hydrate and evaporate to dryness, add hydric sulphate and alcohol, and ignite; if boracic acid be present, the edges of the flame will be coloured green.

A much more beautiful and delicate method of showing the colour which boron imparts to flame, is to mix the dry borates with finely-powdered fluor spar and hydro-potassic sulphate, and to heat this mixture in a loop of platinum wire in a Bunsen flame, the oxidising portion of which will be coloured green.

Questions on Boron and Boracic Acid.

1. How is boron found in nature?

2. How much boron and how much oxygen is there in 21 grammes of boracic acid?—*Ans.* 6.6 grammes of boron and 14.4 grammes of oxygen.

3. How is boracic acid obtained from borax?

4. Describe the behaviour of boracic acid to blue litmus and yellow turmeric paper.

5. In testing for boracic acid why is sodic hydrate added to the solution before evaporating it?

6. What colour does boracic acid impart to flame? and in which part of the flame does the colour appear?

CHAPTER XXVI.

SILICON.

Silicon is found in nature combined with oxygen. This compound is called silicic acid, silica, and, commonly, flint. Silica is found in the form of black flints in chalk formations; it also constitutes a great part of granite. It also occurs crystallised as quartz or rock crystal. Agate and calcedony are amorphous varieties of silica.

Silicon is obtained from a compound of silicon, fluorine and potassium, called potassic silicofluoride. If this substance be heated with potassium, or sodium, and zinc, the temperature not being allowed to rise too high, fluoride of sodium or potassium will be formed; the silicon will dissolve in the melted zinc, and crystallise on cooling. Hydric chloride must be added to dissolve the zinc, and the crystals of silicon will be left. Silicon exists in three allotropic modifications. 1st. Amorphous silicon is a brown powder insoluble in water, and in hydric nitrate or sulphate, but soluble in potassic hydrate, when heated with it, also in hydric fluoride. 2nd. Graphitoid silicon, which, like carbon, exists in hexagonal plates. 3rd. Crystallised silicon, which is crystallised in six-sided prisms.

Amorphous silicon.—Amorphous silicon is formed when potassic silicofluoride is heated with potassium without zinc. It may also be prepared by heating potassium or sodium in a tube through which the vapour of chloride of silicon is passing. If amorphous silicon be heated to a high temperature, its properties are changed, its specific gravity increases, and it is not acted upon by hydric fluoride. In this way amorphous may be converted into graphitoid silicon.

Silicic acid is a compound of silicon and oxygen; 60 grammes of silicic acid contain 28 grammes of silicon, and 2(16) grammes of oxygen. It is a very weak acid, and is easily expelled from its compounds by carbonic acid. In the pure state, when artificially prepared, it is a light, white, insoluble, and very infusible powder. It is insoluble in all acid liquids except hydric fluoride, but it is dissolved by solutions of potassic or sodic hydrate.

When heated to a high temperature with sodic carbonate, carbonic acid is driven off, and sodic silicate remains. If the sodic carbonate be taken in excess, the resulting sodic silicate dissolves slowly in water when cold, but more rapidly when heated; and from this solution the silica can be precipitated by a mixture of ammonia and ammonic chloride.

At high temperatures, silicic acid is able to expel sulphuric acid, whereas at low temperatures less powerful acids are able to replace it in its compounds.

If hydric chloride be added to a dilute solution of a soluble silicate, no apparent change takes place in the liquid; this is owing to the fact that hydric silicate is

soluble to a certain extent in hydric chloride; when, however, hydric chloride is added to a strong solution of silicate of soda, hydric silicate is precipitated, forming a dense, gelatinous mass. When this is dried, excess of chloride will be present; and if this be well washed away, and the residue dried again and ignited, pure silica will be obtained.

We have before considered the diffusion of gases through porous septa, or diaphragms; liquids also have a power of diffusion. Liquids diffuse into one another; for example, a solution of common salt will diffuse into common water. If a vessel containing a solution of common salt be placed carefully, with its mouth covered, in a vessel containing water, the water being sufficiently deep to cover the vessel of salt and water, and if the cover be removed from that vessel, in time the salt and water solution will diffuse out into the larger vessel and the water into the smaller vessel, until both liquids are of equal density.

Professor Graham, who has most carefully investigated this subject, finds that certain bodies, which he terms colloid, will not pass through porous septa; whereas crystalline substances in solution will do so. Hydric silicate is a colloid body. Now, if to a dilute solution of an alkaline silicate excess of hydric chloride be added, and if these be placed in what Professor Graham calls a dialyzer, the solution of sodic chloride will diffuse out, and leave behind the hydric silicate in aqueous solution. Such a solution cannot be made which contains more than 5 per cent. of silicic acid, for stronger solutions than this almost immediately gelatinise; in fact, after a time, even a dilute

solution gelatinises, and the presence of foreign bodies, such as dust, expedite this.*

Silicic chloride is a colourless liquid, very volatile, and immediately decomposed into silica and hydric chloride when it comes in contact with water or moisture; therefore in the preparation of this body, all moisture must be carefully got rid of. Silicic chloride boils at 50°C.

To make silicic chloride, silica should be mixed with finely-powdered charcoal, and be made into pellets with oil or sugar and water; these should be heated in a closed crucible; by this means the oil or sugar is reduced to carbon, which is in very close contact with the particles of silica. The pellets should be put into a porcelain tube and heated in a gas furnace similar to that employed in the analysis of atmospheric air, and through the tube a current of carefully-dried chlorine should be passed; the vapour of silicic chloride will pass out at the other end of the tube; it should be led into a U tube, surrounded by a mixture of ice and salt; in the curved part of the U tube a piece of ordinary glass tubing should be fastened, so that the condensed silicic chloride may, as it is formed, run into a bottle, fitted to the straight tube to receive it.

Glass.—Silica is employed in the manfacture of all kinds of glass. Sand is the form in which the silica is used. There are several kinds of glass. Flint glass, which is used for tumblers, wine glasses, &c.:

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* *Silicates.*—It will be more convenient to treat of these compounds in Part II., as also of the compounds of silicon with fluorine.

this kind contains silica (pure white sand), oxide of lead, and potash; it is a double silicate of lead and potash. Common window-glass contains sand, lime, and soda, it is therefore a double silicate of lime and soda. Bohemian glass contains sand, potash, and lime; and plate glass contains sand, lime, and soda, purer materials being used in the manufacture of plate glass than in that of crown or common window glass. A finer kind of flint glass is used for optical purposes. These different kinds of glass are made in different ways, and other substances are employed in their manufacture besides those already mentioned, but which do not enter into their permanent composition.

These glasses are silicates, having different and perhaps definite compositions. They combine with metallic oxides and derive from them various beautiful colours. Ruby glass is coloured with oxide of gold; some kinds with cuprous oxide; green glass is coloured with cupric oxide, also with chromic oxide, and sometimes with ferrous oxide. Blue glass receives a beautiful purplish blue colour from oxide of cobalt; white glass can be superficially stained yellow by oxide of silver.

Manganese, which gives a purple colour to glass, is used, under the name of glass-maker's soap, for counteracting the green colour which iron, an impurity in sand, always imparts to glass when in the form of ferrous oxide. It appears that manganic binoxide keeps the iron oxidised to the state of ferric oxide during the manufacture of the glass, and this oxide gives it a yellow colour, but inasmuch as it requires large quantities to produce a perceptible tint, and

as only small quantities of iron are present, the manganic binoxide almost entirely destroys the colouring power of the iron oxide. Manganese stains glass purple when it is in excess, and this is why so much plate glass of a purple tint is to be seen in windows.

Of late years, a silicate has been manufactured which is soluble in water; it is applied to various purposes, and has various names—soluble glass, water glass, &c. It is a silicate of soda or potash, in which the alkali is in excess. When dissolved and concentrated it is of the consistence of treacle, colourless when pure, but generally of a yellowish tint; it is very sticky, and might well be used for gum; it is employed in soap-making, in the manufacture of artificial stone, for the preservation of stone, and in wall painting. In the three last cases it is supposed to form insoluble double silicates with the substances with which it is mixed, or to which it is applied.

Tests for Silica.

A mixture of ammonia and ammonic chloride precipitates silica as a white hydrate from solutions containing it when it is in tolerably large quantities; minute traces are not readily precipitated in this way. The only satisfactory test for silica is to evaporate the silicate to dryness in the presence of hydric chloride, and to dry the residue thoroughly, so as to expel all moisture. All that is soluble should be dissolved with water, and the silica which remains will be found to be insoluble in hydric nitrate or chloride.

Questions on Silicon and its Compounds.

1. In what forms can silicon exist?
2. How can crystallised silicon be prepared?
3. To what other substance is graphitoid silicon like?
4. How does silicic acid occur in nature?
5. Give some reasons for calling silica an acid.
6. What is meant by a colloid body? Describe Graham's method of dialysis.
7. How much silicon and how much oxygen is there in 7.8 grammes of silicic acid?—*Ans.* 3.64 grammes of silicon and 4.16 grammes of oxygen.
8. How is silicic chloride prepared? What precaution must be used in its preparation?
9. What is the constant constituent of all kinds of glass in use?
10. State the difference between window glass and flint glass.
11. Why is manganese used in the manufacture of glass?
12. How can a soluble glass be made?
13. Describe clearly the method of testing for silica.

PART II.

SYMBOLS, NOTATION, AND THEORETICAL CONSIDERATIONS.

CHAPTER I.

In the first part of this book the weights of all elements and compounds have been taken in grammes, and the reason why they were so taken is, that it is far easier to understand relative proportions when definite weights are taken, than when abstract numbers are employed.

If the previous chapters have been carefully studied, it will have been noticed that in all compounds no less weight of oxygen than 16 grammes has existed; in several cases, however, more than 16 grammes exist in bodies which have been treated of; for example, in 98 grammes of hydric sulphate there are 64 grammes of oxygen; now 64 is 4 times 16, and in the later chapters the quantity of oxygen present in a compound has been written thus—4(16) grammes; and this has been pretty generally done with respect to other elements when they have been present in different proportions, the larger quantities being indicated in the same way. This course has been adopted to accustom the student to connect the idea of certain numbers

with particular elements. Thus, when oxygen is mentioned, the mind should connect with it the idea of 16, when sulphur 32, when hydrogen 1, and so on. It was partly for this purpose that the working of simple calculations was advised.

If we take the *lowest* number of grammes in which any element enters into combination—take for instance, 14 grammes of nitrogen—we shall find that nitrogen always combines in this proportion, or in multiples of it—as in 108 grammes of nitric acid there are 2(14) grammes of nitrogen, but in 63 of hydric nitrate only 14; in 142 grammes of phosphoric acid there are 2(31) grammes of phosphorus, but in 98 grammes of trihydric phosphate there are only 31 grammes of phosphorus.

It will not now be difficult to understand that if we take the weight of any compound in grammes, the *proportions* in which the elements forming it exist in that compound will be just the same if we take any other weight of it, be that weight large or excessively small. Thus, there are 4(16) grains of oxygen in 98 grains of hydric sulphate, or 4(16) tons in 98 tons of it.

From what has been stated, then, we learn that with hydrogen (the lightest substance in nature) other elements enter into combination in certain definite proportions by weight, and that these proportions, as far as we know, vary only in this, that they may be multiples of the lowest weight of an element which can combine with hydrogen, and that elements enter into combination with one another in the proportionate weights in which they enter into combination with

hydrogen. Hydrogen, then, is taken as the standard to which other elements are referred, and its weight is assumed to be 1. Thus 35.5 parts of chlorine combine with 1 part of hydrogen forming hydric chloride, but, if potassium act on hydric chloride, it is found that 1 part of hydrogen goes away, and 39 parts of potassium take its place; so that 39 parts of potassium take the place of 1 of hydrogen, and combine with 35.5 of chlorine. But if potassium act on water, which consists of 2 parts of hydrogen and 16 parts of oxygen, then 39 parts of potassium take the place of 1 of hydrogen, so that 39 parts of potassium can enter into combination with 16 parts of oxygen and 1 part of hydrogen which remain in the original water.

In such cases as these, where a definite weight of an element can replace a definite weight of another element, that weight is called an "equivalent," that is, it is equal in its power of replacing. 39 parts of potassium are equivalent to 1 part of hydrogen and to 35.5 of chlorine. 2 parts of hydrogen unite with 16 parts of oxygen to form water, therefore 8 parts of oxygen are "*equivalent*" to 1 part of hydrogen, though not less than 16 parts of oxygen can unite chemically with hydrogen. And as 39 parts of potassium are equivalent to 1 of hydrogen, they are equivalent to 8 of oxygen. An equivalent, therefore, may be defined to be, *that weight of an element or compound which is capable of replacing a definite weight of any other element or compound.*

CHAPTER II.

ATOMIC THEORY.

The atomic theory was devised by Dalton. It supposes that elements can be divided and subdivided until they can be divided no further. If this assumption be true, it is clear that an excessively small particle must be left which cannot be divided, and this small particle is called an "atom"; but this atom, it is asserted, in most cases does not exist alone, or in the free state; it is generally combined with another atom, either of the same or some other element. Now, to these atoms of different elements, weights have been assigned, and those weights are no others than those which we have already been using in grammes. A list of the atomic weights of the different elements will be given at the end of this chapter.

An atom is defined to be the smallest quantity of an element which can exist in combination. Now if an atom admits of no further division, a less quantity than an atom cannot enter into combination with other elements. The weight of the atom of hydrogen is assumed to be 1; and as we have seen the relative weights of equal

volumes, *i.e.* 11.2 litres of hydrogen and oxygen are 1 and 16, therefore the weight of the atom of oxygen is 16 if that of hydrogen be 1; and as water is found to contain 1 volume of oxygen and 2 volumes of hydrogen, we conclude that there are 2 atoms of hydrogen and 1 atom of oxygen in water, and therefore that 1 part of hydrogen is equivalent to 8 parts of oxygen. You see, then, that the atomic weight of oxygen is 16, but its equivalent weight with respect to hydrogen is 8, therefore atomic weight and equivalent weight are not in this case the same thing. There are other cases in which the atomic and equivalent weights are different. A rule for finding equivalent weights will be given presently.

Valency.—The atom of oxygen combines with not less than 2 atoms of hydrogen, or with 2 atoms of potassium, or of sodium, and certain other elements; but 1 atom of hydrogen combines with 1 atom of chlorine, and 1 atom of chlorine combines with 1 of potassium. It appears, then, that oxygen has this different property from these other elements, that whereas they can combine with one another, atom for atom, oxygen requires 2 atoms of either of them, or 1 of each of any two of them, to form a combination.

This property was formerly called "atomicity," it is now termed "valency," from "*valeo.*"

The word "atomicity" was, however, used in a somewhat different sense by Kekulè, and some other chemists. Hydrogen, and those elements with which it can combine atom for atom, are called *monovalent.* Oxygen, and elements like it, which combine with not less than 2 atoms of hydrogen, or other monovalent elements, are called "*divalent.*" Nitrogen, and ele-

ments like nitrogen, which combine with 3 atoms of hydrogen or other monovalent elements, are called "*trivalent.*" Carbon, and elements like carbon which combine with 4 atoms of a monovalent element, are called tetravalent.

In the list of atomic weights given at the end of this chapter, the elements will be classed under two heads, viz. those of even and those of uneven valency. It will be necessary for the student to make himself acquainted with the valency of those elements which he most commonly meets with: also with their atomic weights.

Oxygen is divalent. If we divide its atomic weight by 2, we get 8, *i.e.* $\frac{16}{2} = 8$, this is its "equivalent" with respect to monovalent elements. Nitrogen is trivalent. If we divide its atomic weight 14 by 3, we get 4.66, this is its equivalent with respect to hydrogen, and other monovalent elements. Carbon is tetravalent. Its atomic weight is 12, and if we divide this by 4, we have 3 as the equivalent of carbon with respect to hydrogen, or a monovalent element.

Molecules.—The molecule is the least quantity of any substance which can exist in a free state, and is generally composed of two or more atoms; thus the least quantity of hydrogen which can exist in a free state, *i.e.* the molecule, consists of 2 atoms. The smallest possible quantity of water which can exist in the free state consists of 3 atoms, namely, 2 of hydrogen and 1 of oxygen. The smallest quantity of hydric sulphate consists of 7 atoms, 2 of hydrogen each weighing 1, 1 of sulphur weighing 32, and 4 of oxygen each weighing 16. The molecule of hydric

nitrate contains 5 atoms: 1 of hydrogen, 1 of nitrogen weighing 14, and 3 of oxygen weighing 16 each. We know that this is the molecule of hydric nitrate, and not double this quantity, because in it hydrogen can be replaced by a monovalent element only in one proportion.

When we took the weights of compounds in grammes we found that, without exception, those weights of the compounds we considered, in a state of vapour occupied 2(11.2) litres.

Professor Williamson, who suggested the use of this measure, namely, 2(11.2) litres, or 2 volumes, calls the weight of the element or compound occupying it the *absolute molecule*. We shall revert briefly to this subject after treating of symbols and formulæ.

Table of Atomic Weights.*

ELEMENTS OF UNEVEN VALENCY.

Hydrogen	H = 1	Silver	Ag = 108
Chlorine	Cl = 35.5	Nitrogen	N = 14
Bromine	Br = 80	Phosphorus	P = 31
Iodine	I = 127	Arsenic	As = 75
Fluorine	F = 19	Antimony	Sb = 122
Sodium	Na = 23	Bismuth	Bi = 210
Potassium	K = 39	Boron	B = 11
Lithium	Li = 7	Gold	Au = 196

* Some elements have two equivalents, they will be noticed at the end of Part II.

ELEMENTS OF EVEN VALENCY.

Oxygen	O	=	16	Iron	Fe	=	56
Sulphur	S	=	32	Chromium	Cr	=	52.5
Calcium	Ca	=	40	Manganese	Mn	=	55
Strontium	Sr	=	87.5	Cobalt	Co	=	58.5
Barium	Ba	=	137	Nickel	Ni	=	58.5
Lead	Pb	=	207	Copper	Cu	=	63.5
Mercury	Hg	=	200	Carbon	C	=	12
Magnesium	Hg	=	24	Silicon	Si	=	28
Zinc	Zn	=	65	Platinum	Pt	=	197
Cadmium	Cd	=	112	Tin	Sn	=	111
Aluminum	Al	=	27	Molybdenum	Mo	=	98

CHAPTER III.

SYMBOLS AND FORMULÆ.

Symbols are used to denote elements, or rather the elements taken in the proportions of their atomic weights. H does not represent the element hydrogen, but a certain weight of hydrogen ; and as we have assumed the weight of hydrogen to be 1, the symbol H always means 1 part of hydrogen by weight. O is the symbol for 16 parts by weight of oxygen. The symbol for water (which contains 2 parts by weight of hydrogen and 16 parts by weight of oxygen) will be H_2O. S, the symbol for sulphur, denotes 32 parts by weight of it. Now, as in hydric sulphate, there are 2 parts by weight of hydrogen, 32 of sulphur, and 4(16) of oxygen, the formula for it is H_2SO_4. N equals 14 parts by weight of nitrogen; and, as in hydric nitrate, there are 1 part of hydrogen, 14 of nitrogen, and 3(16) of oxygen, its formula will be HNO_3. In ammonia there are 14 parts by weight of nitrogen, and 3 of hydrogen, its formula therefore will be NH_3. C equals 12 parts by weight of carbon. In carbonic acid there are 12 parts of carbon and 2(16) of oxygen, and its formula is CO_2; whereas the formula of carbonic oxide (which contains 12 parts of carbon to 16 of oxygen) is CO ; and that of marsh gas, which contains 4 parts of hydrogen and 12 of carbon, is CH_4.

We need not give any more illustrations, as the

symbols of all the elements referred to in this work and some others, are given in the table of atomic weights; practice, however, will be necessary to render the student familiar with them.

FORMULÆ FOR REACTIONS GIVEN IN PART I.

Molecular formulæ are employed, *i.e.* when an element is represented as *free*, in an equation, the molecule and not the atom is given, or it may be some multiple of the molecule. There are, however, some elements the molecule of which is not known, and others whose molecule consists of more or less than two atoms; such elements, when represented in the free state, are marked with an asterisk, thus $\overset{*}{Cu}$.

HYDROGEN.

[*See page 11, Part I.*]

Symbol H; molecular weight $H_2 = 2 = 2$ volumes.

Formation of hydrogen by the action of zinc on hydric sulphate—

$$\overset{*}{Zn} + H_2SO_4 = ZnSO_4 + H_2.$$

In this reaction it is seen that SO_4 unites as a whole with Zn, and that H_2 is set free. It may here be noticed that if H_2SO_4 react with ZnO (zinc oxide), the hydrogen is not set free, for

$$ZnO + H_2SO_4 = ZnSO_4 + H_2O.$$

Here water is formed, and this is the case generally when the hydrogen salts of acids act upon metallic oxides.

Formation of hydrogen by the action of zinc on potassic hydrate—

$$\overset{*}{Zn} + 2(KHO) = K_2ZnO_2 + H_2.$$

In this case zinc expels hydrogen, and potassic zincate is formed. ZnO_2 behaves to K_2 as SO_4 did in in the former case to Zn. When steam is passed over red-hot iron filings, the reaction which takes place is as follows :—

$$Fe_3 + 4(H_2O) = Fe_3O_4 + 4(H_2).$$

Fe (the abbreviation for ferrum) is the symbol for iron. Fe_3O_4 is called magnetic oxide of iron. It is supposed to be formed of two oxides, viz. FeO, ferrous oxide, and Fe_2O_3, which is usually called ferric oxide. The colour of magnetic oxide of iron is black, that of Fe_2O_3 (iron rust) is red.

When water is acted upon by sodium, the equation for the reaction is

$$Na_2 + 2(H_2O) = 2(NaHO) + H_2.$$

Na is the abbreviation for natrium, sodium ; here sodic hydrate is formed and hydrogen set free. If potassium be used instead of sodium,

$$K_2 + 2(H_2O) = 2(KHO) + H_2.$$

K is the initial letter of kalium, potassium. In the action of zinc on hydric sulphate, 1 atom of zinc took the place of 2 atoms of hydrogen, because zinc is a divalent element. In the reaction between zinc and potassic hydrate, it was necessary to take 2 molecules of potassic hydrate (KHO is a molecule), for zinc being divalent could not replace *only 1* atom of hydrogen.

OXYGEN.

[*See page 18, Part I.*]

Symbol O = 16 ; molecular weight, O_2 = 32 = 2 volumes.

Formation of oxygen from potassic chlorate—

$$2(KClO_3) = 2KCl + 3O_2.$$

2 molecules of potassic chlorate, $KClO_3$, are here taken because, in the free state, the molecule of oxygen is composed of 2 atoms, that is O_2, therefore free oxygen must contain an even number of atoms. When manganic binoxide is heated to a high temperature the decomposition is

$$3(MnO_2) = Mn_3O_4 + O_2.$$

Mn is the symbol for the atom of manganese, MnO_2 that for the binoxide; binoxide meaning an oxide containing 2 atoms of oxygen; the molecule of the binoxide is taken three times because only one-third of its oxygen can be obtained from manganic binoxide by heat; and as the molecule of oxygen in the free state consists of 2 atoms, no less a quantity than 3 molecules of the binoxide would furnish it. The residue Mn_3O_4 resembles the iron oxide, whose formula is Fe_3O_4, in its composition.

Other substances are used for the production of oxygen which were not mentioned in Part I. on account of the more complicated nature of their reactions, for example, when hydric sulphate acts upon potassic dichromate oxygen is evolved.

$$\underset{\text{(Potassic dichromate.)}}{2(K_2Cr_2O_7)} + 8(H_2SO_4) = 2(K_2SO_4)$$

$$+ \underset{\text{(Chromic sulphate.)}}{2\{Cr_2(SO_4)_3\}} + 8(H_2O) + 3O_2.$$

Cr is the symbol for the atom of chromium.

When mercury is heated in air or oxygen it becomes mercuric oxide, symbol HgO. When mercuric oxide is heated to a higher temperature the following decomposition takes place:—

$$2(HgO) = 2\overset{*}{Hg} + O_2.$$

A similar action takes place with barytic oxide. If barytic oxide, whose molecule is BaO, be heated in air, it takes up oxygen and becomes binoxide, BaO_2 but an increase of temperature expels the second atom of oxygen—

$$2(BaO_2) = 2(BaO) + O_2.$$

NITROGEN.

[*See page 25, Part I.*]

Symbol N = 14; molecular weight, $N_2 = 28 = 2$ volumes.

When chlorine gas acts upon ammonia, and nitrogen is evolved, the equation representing the reaction is as follows:—

$$2(NH_3) + 3Cl_2 = 6(HCl) + N_2.$$

If ammonic nitrite be heated, nitrogen is evolved; thus—

$$NH_4 NO_2 = 2(H_2O) + N_2.$$

Ammonic nitrite is hydric nitrite in which the hydrogen has been replaced by ammonium, NH_4. Nitrogen can be obtained by boiling potassic nitrite with ammonic chloride—

$$KNO_2 + NH_4Cl = KCl + 2(H_2O) + N_2.$$

WATER.

[*See page 39, Part I.*]

Symbol H_2O; molecular weight 18, $H_2 = 2$, $O = 16$, $H_2O = 2$ volumes; density 9.

When hydrogen is passed over red-hot cupric oxide, the reaction is represented thus—

$$CuO + H_2 = \overset{*}{Cu} + H_2O.$$

And when over ferric oxide, thus—

$$Fe_2O_3 + 3\,H_2 = \overset{*}{Fe_2} + 3(H_2O).$$

CHAPTER IV.

CARBONIC ACID.

[*See page 71, Part I.*]

Symbol CO_2, molecular weight 44; $C = 12, O_2 = 32$, therefore $CO_2 = 44$. CO_2 occupies 2 volumes, therefore its density is 22.

Carbonic acid is usually obtained from marble; but marble, chalk, Iceland spar, limestone, all have the same chemical composition, and if treated in the same way will yield carbonic acid. The symbol of either of them is $CaCO_3$, *i.e.* CaO calcic oxide, and CO_2 carbonic acid.

$$CaCO_3 + \underset{\text{Hydric chloride.}}{2(HCl)} = \underset{\text{Calcic chloride.}}{CaCl_2} + H_2O + CO_2.$$

2 molecules of hydric chloride are here taken, and this is because lime is divalent and chlorine is monovalent; and so 2 molecules of hydric chloride are necessary to yield 2 atoms of chlorine, which lime requires to form calcic chloride. Carbonic acid is set free just the same when calcic carbonate is acted upon by hydric sulphate, but in this case calcic sulphate which is formed is insoluble, and adhering to the marble pre-

vents the continued action of the hydric sulphate. If, however, hydric sulphate be employed, the equation will be—

$$CaCO_3 + H_2SO_4 = CaSO_4 + H_2O + CO_2.$$

In this case only 1 molecule of H_2SO_4 is required, as it contains 2 atoms of hydrogen which can be replaced by calcium. Because the molecule of hydric chloride contains only 1 atom of hydrogen which can be replaced by a metal, it is called monobasic, as the element which can replace that atom of hydrogen is called monovalent.

Hydric sulphate, which contains 2 atoms of replaceable hydrogen, is called dibasic, as calcium and mercury, which can replace those hydrogen atoms, are called divalent. The basicity of other hydrogen salts will be noticed as occasion may require.

CARBONIC OXIDE.

[*See page 80, Part I.*]

Symbol CO, molecular weight 28; C = 12, O = 16, CO = 28. CO occupies 2 volumes, its density therefore is 14.

When carbonic acid is passed over hot charcoal, carbonic oxide is formed, thus—

$$CO_2 + \overset{*}{C} = 2\,CO.$$

1 molecule of carbonic acid and 1 atom of carbon give 2 molecules of carbonic oxide. CO_2 occupies 2 volumes; but when it unites with an atom of carbon, C, the product, carbonic oxide, 2CO, occupies 4 volumes. Carbonic oxide is prepared in the laboratory by heating hydric oxalate with hydric sulphate. The molecular formula of hydric oxalate is $C_2H_2O_4$; or it may be written

thus—CO_2COH_2O, that is, carbonic acid, carbonic oxide, and water. The action of hydric sulphate is to break it up in this way; the hydric sulphate is not decomposed, therefore it will appear unchanged on both sides of the equation

$$C_2H_2O_4 + H_2SO_4 = CO_2 + CO + H_2SO_4,\ H_2O.$$

In order to get the carbonic oxide by itself, the gases must be passed through a solution of sodic hydrate, with which the carbonic acid unites to form sodic carbonate, and the carbonic oxide passes on. Another method of obtaining this gas is from hydric formiate: its molecular formula is CH_2O_2, or written thus—COH_2O. It will be seen that if the elements of water be taken from it, carbonic oxide will be set free. In this case, as in the last, hydric sulphate is used with application of heat—

$$CH_2O_2 + H_2SO_4 = CO + H_2SO_4,\ H_2O.$$

When carbonic oxide is heated for a long time with potassic hydrate, potassic formiate is produced—

$$CO + KHO = KCHO_2.$$

The molecule of hydric formiate has 2 atoms of hydrogen, one only can be replaced by a metal. In this respect it is like hydric nitrate, it is therefore called monobasic.

MARSH GAS.

[*See page 83, Part I.*]

Symbol CH_4; molecular weight 16, $C = 12$, $H_4 = 4$, $CH_4 = 2$ volumes; therefore its density is 8.

When marsh gas is made from sodic acetate and soda lime, the reaction is as follows—

(Sodic acetate.)

$$NaC_2H_3O_2 + NaHO = Na_2CO_3 + CH_4.$$

The lime, as stated in Part I., is for the most part unchanged, though some may be changed into carbonate. A molecule of marsh gas requires for its complete combustion 2 molecules of oxygen.

$$\underset{2\text{ vols.}}{CH_4} + \underset{4\text{ vols.}}{2\ O_2} = \underset{2\text{ vols.}}{CO_2} + \underset{4\text{ vols.}}{2(H_2O)}$$

OLEFIANT GAS.

[*See page 90, Part I.*]

Symbol C_2H_4; molecular weight 28, $C_2 = 24$, $H = 4$, $C_2H_4 = 2$ volumes; therefore its density is 14.

Olefiant gas is made from alcohol and hydric sulphate. The molecular formula for alcohol is C_2H_6O; or it may be written thus—$C_2H_4H_2O$.

$$C_2H_6O = C_2H_4 + H_2O.$$

The hydric sulphate need not be expressed in the equation, for that, which acts in the decomposition on the alcohol in forming olefiant gas, is unchanged. It has already been noticed that a part of the alcohol used is broken up into other compounds besides olefiant gas.

Equal measures of chlorine and olefiant gas form Dutch liquid; its molecular formula, therefore, is $C_2H_4Cl_2$. There is here a condensation of volume. C_2H_4 occupies 2 volumes, and Cl_2 occupies 2 volumes; but when combined, in the state of vapour $C_2H_4Cl_2$, occupies 2 volumes. A molecule of olefiant gas requires for its complete combustion 3 molecules of oxygen.

$$\underset{2\text{ vols.}}{C_2H_4} + \underset{6\text{ vols.}}{3O_2} = \underset{4\text{ vols.}}{2(CO_2)} + \underset{4\text{ vols.}}{2(H_2O)}$$

CHAPTER V.

OXIDES OF NITROGEN.

The formulæ of the oxides of nitrogen, which are five in number, are as follow:—

Nitrous oxide N_2O,
Nitric oxide NO,
Nitrous acid, N_2O_3,
Nitric peroxide N_2O_4,
Nitric acid N_2O_5.

These compounds of oxygen with nitrogen illustrate well the law of *multiple proportions*. In N_2O, 2 atoms of nitrogen, each = 14, combine with 1 atom of oxygen = 16. In NO, 1 atom of nitrogen = 14, combines with 1 atom of oxygen = 16. In nitrous acid N_2O_3, 2 atoms of nitrogen, each = 14, combine with 3 of oxygen, each = 16. In peroxide of nitrogen, N_2O_4, 2 atoms of nitrogen, each = 14, combine with 4 atoms of oxygen, each = 16; and in nitric acid 2 atoms of nitrogen, each = 14, combine with 5 atoms of oxygen, each = 16. From these observed facts Dalton founded the law of multiple proportions, viz. *that elements combine in the proportion of their atomic weights, or in multiples of those atomic weights.* There is no instance of an element combining in the propor-

tion of a fraction of its atomic weight; such a body does not exist as $N\frac{O}{2}$, or $N = 14$, $\frac{O}{2} = 8$.

NITROUS OXIDE.

[*See page 99, Part I.*]

Symbol N_2O; molecular weight 44, $N_2 = 28$; $O = 16$, $N_2O = 2$ volumes; therefore its density is 22.

Nitrous oxide is prepared by heating ammonic nitrate. The decomposition is represented thus—

(Ammonic nitrate.)

$$NH_4NO_3 = 2(H_2O) + N_2O.$$

NITRIC OXIDE.

[*See page 101, Part I.*]

Symbol NO; molecular weight 30, $N = 14$, $O = 16$, $NO = 2$ volumes; therefore its density is 15.

When copper is mixed with hydric nitrate of about specific gravity 1.2, nitric oxide is evolved, thus—

(Cupric nitrate.)

$$3\overset{*}{Cu} + 8(HNO_3) = 3\{Cu(NO_3)_2\} + 4(H_2O) + 2(NO).$$

3 atoms of copper replace 6 atoms of hydrogen, because copper is divalent. When nitric oxide is mixed with hydrogen, and passed through a red-hot tube containing spongy platinum, ammonia and water are formed, thus—

$$NO + 5H = NH_3 + H_2O.$$—*Williamson.*

Potassium burns in nitric oxide; it combines with the oxygen, and nitrogen is left. Phosphorus also burns in it almost as brilliantly as in oxygen. Nitric oxide dissolves in ferrous sulphate, forming a brown

liquid; the colour, however, is dispelled by a gentle heat. This property is made use of as a test for a nitrate.

The molecule of this gas is sometimes given as N_2O_2; in this case it occupies 4 volumes.

NITROUS ACID.

[*See page 102, Part I.*]

Symbol N_2O_3; molecular weight 76, $N_2 = 28$, $O_3 = 48$, $N_2O_3 = 2$ volumes; therefore its density is 38.

When nitric oxide comes in contact with free oxygen, nitrous acid is formed

$$4(NO) + O_2 = 2(N_2O_3).$$

Nitrous acid is a red gas; when it is shaken up with water, the colour disappears, and hydric nitrate and nitric oxide are formed—

$$3(N_2O_3) + H_2O = 2(HNO_3) + 4(NO).$$

Nitrous acid combines with oxides, forming salts; thus nitrous acid and potassic oxide give potassic nitrite—

$$N_2O_3 + K_2O = K_2N_2O_4.$$

Here the molecule is half this quantity,

$$\frac{K_2N_2O_4}{2} = KNO_2,$$

as only one nitrite of potassium, which contains no hydrogen, exists.

Hydric nitrite has not been obtained.

NITRIC PEROXIDE.

[*See page 103, Part I.*]

Symbol N_2O_4; molecular weight 92, $N_2 = 28$, $O_4 = 64$, $N_2O_4 = 2$ volumes; therefore its density is 46.

When plumbic nitrate is heated, nitric peroxide and oxygen are set free—

$$\underset{\text{(Plumbic nitrate.)}}{2\{Pb(NO_3)_2\}} = \underset{\text{(Plumbic oxide.)}}{2(PbO)} + 2(N_2O_4) + O_2.$$

Lead is divalent, therefore it replaces 2 atoms of hydrogen in hydric nitrate, thus—

$$Pb \begin{array}{l|l} H & NO_3 \\ H & NO_3 \end{array}$$

NITRIC ACID.

[*See page 104, Part I.*]

Symbol N_2O_5; molecular weight 108, $N_2 = 28$, $O_5 = 80$.

Chlorine gas passed over dried crystals of argentic nitrate, heated to about 90°C., sets free nitric acid and oxygen—

$$4(AgNO_3) + 2Cl_2 = 4(AgCl) + 2(N_2O_5) + O_2.$$

When the vapour of nitric acid is passed over strongly-heated metallic copper, the following equation expresses the reaction:—

$$N_2O_5 + 5\overset{*}{Cu} = 5(CuO) + N_2.$$

HYDRIC NITRATE.

[*See page 105, Part I.*]

Symbol HNO_3; molecular weight 63, $H = 1$, $N = 14$, $O_3 = 48$.

When potassic nitrate and hydric sulphate are heated together in a retort, hydric nitrate distils over.

$$KNO_3 + H_2SO_4 = KHSO_4 + HNO_3.$$

This equation represents the action at such a temperature as can be used with a glass retort. The substance $KHSO_4$ is hydro-potassic sulphate; when this is

heated with another molecule of potassic nitrate to a higher temperature a further action takes place.

$$KNO_3 + KHSO_4 = K_2SO_4 + HNO_3,$$

another molecule of hydric nitrate being formed, and K_2SO_4 potassic sulphate being left.

The formula, therefore, for the complete decomposition of hydric sulphate by potassic nitrate is

$$2(KNO_3) + H_2SO_4 = K_2SO_4 + 2(HNO_3)$$

When nitric acid meets with water, it unites with it energetically.

$$N_2O_5 + H_2O = H_2N_2O_6.$$

Why is not this the molecule of hydric nitrate? Because the molecule of hydric nitrate is found to have only 1 atom of hydrogen which can be replaced by a metal, and that hydric nitrate contains no hydrogen but replaceable hydrogen, therefore—

$$\frac{H_2N_2O_6}{2} = HNO_3$$

is the molecule.

CHAPTER VI.

THE WATER TYPE.

A very clear and rational method of regarding certain compounds and their decompositions was originated by Professor Williamson; it was afterwards adopted by many eminent chemists. It consists in regarding these bodies as formed on the type of one, two, or more molecules of water. The formula for water, H_2O, may be written thus—

$$\begin{matrix}H\\H\end{matrix}O.$$

When potassium acts on water, as you know, potassic hydrate, KHO, is formed by the displacement of 1 atom of hydrogen of the water by potassium, thus—

$$\begin{matrix}K \mid H\\ \\ H\end{matrix}O\text{, it is then written }\begin{matrix}K\\H\end{matrix}O\text{, the hydrogen}$$

having been liberated. When bodies are regarded in this way they are said to be represented on the type of a molecule of water and this method of representation is called the "water type."

Hydric nitrate may be expressed in this way, thus—

$$\begin{matrix} & H \\ NO_2 \mid & H\end{matrix}O$$

And NO_2, which replaces the atom of hydrogen, is called a radical, the radical of nitric acid; for if the other atom of hydrogen be removed by another atom of NO_2 we get this body—

$$\left.\begin{array}{l} NO_2 \mid H \\ NO_2 \mid H \end{array}\right\} O = N_2O_5$$

Potassic nitrate is a body formed on this type, in which an atom of hydrogen has been replaced by potassium, and the other by the radical NO_2.

$$\left.\begin{array}{l} K \mid H \\ NO_2 \mid H \end{array}\right\} O = KNO_3$$

Hydric sulphate and other sulphates are written on the type of 2 molecules of water; now remember that hydric sulphate is called dibasic,

$$\left.\begin{array}{l} SO_2 \mid H^2 \\ \quad H_2 \end{array}\right\} O_2 = H_2SO_4$$

$$\left.\begin{array}{l} SO_2 \mid H_2 \\ K_2 \mid H_2 \end{array}\right\} O_2 = K_2SO_4$$

SO_2 being the radical of sulphates.

Trihydric phosphate is called tribasic; all its three

atoms of hydrogen are replaceable by metals; it is represented on the type of 3 molecules of water:—

$$\left.\begin{array}{l} PO \mid H_3 \\ \quad H_3 \end{array}\right\} O_3 = H_3PO_4$$

and

$$\left.\begin{array}{l} PO \mid H_3 \\ Na_3 \mid H_3 \end{array}\right\} O_3 = Na_3PO_4$$ trisodic phosphate, or sodic phosphate.

PO is the radical of the phosphates.

AMMONIA.

[*See page 113, Part I.*]

Symbol NH_3; molecular weight 17; $N = 14$, $H_3 = 3$, $NH_3 = 2$ vols.; density 8.5.

Ammonic chloride, when mixed with calcic oxide, is decomposed, and ammonia is evolved—the decomposition begins in the cold, but heat is required to complete it.

$$\underset{\text{(Ammonic chloride.)}}{2(NH_4Cl)} + \underset{\text{(Calcic oxide.)}}{CaO} = \underset{\text{(Calcic chloride.)}}{CaCl_2} + H_2O + 2(NH_3)$$

In this reaction 2 molecules of ammonic chloride are taken, because calcium, being divalent, requires 2 atoms of monovalent chlorine to unite with it.

Ammonic chloride is formed by the direct union of ammonia NH_3 with hydric chloride HCl, and is sometimes written thus—NH_3HCl, and is regarded as a hydrochlorate of ammonia; when written NH_4Cl it is considered to be a chloride of ammonium.

NH_4 is looked upon by chemists as a metal, and is called ammonium. If a compound of mercury and sodium, termed sodium amalgam, be added to a saturated solution of ammonic chloride,—the chlorine combines with the sodium, forming sodic chloride, and a new amalgam of ammonium and mercury is produced. This amalgam is extremely bulky, and is very easily broken up into free hydrogen, ammonia, and mercury. The hydrogen and ammonia occupying those volumes in which they form ammonium, viz. 1 volume of hydrogen and 2 of ammonia gas—H, NH_3.

When ammonia unites with water, hydrated oxide of ammonium (ammonic hydrate) is formed—just as potassium forms with water potassic hydrate.

Potassic hydrate.	Ammonic hydrate.
$\left.\begin{matrix} K \\ H \end{matrix}\right\} O.$	$\left.\begin{matrix} NH^4 \\ H \end{matrix}\right\} O.$

When hydric chloride is added to a solution of ammonic hydrate a decomposition occurs, thus—

$$\begin{array}{c|c} NH_4 & HO \\ Cl & H \end{array}$$

And when hydric nitrate reacts with ammonia

$$\begin{array}{c|c} NH_4 & HO \\ NO_3 & H \end{array} \qquad \begin{array}{c|c} (NH_4)_2 & (HO)_2 \\ SO_4 & H_2 \end{array}$$

In the first case, the chlorine of the hydric chloride unites with the ammonia of the ammonic hydrate, and its hydrogen combines with HO to form water; and in the second, NO_3 or SO_4 unite with NH_4 to form NH_4 NO_3 or $(NH_4)_2$ SO_4 ammonic nitrate or ammonic sulphate, and H with HO to form H_2O water.

When hydric nitrate acts upon zinc, the nitrate being

of specific gravity about 1.2, nitric oxide is produced but "when zinc or tin is used with a stronger acid, ammonia is amongst the products ; for instance—

$$4\overset{*}{Zn} + 9HNO_3 = 4\{Zn(NO_3)_2\} + 3(H_2O) + NH_3,$$

the ammonia combining with excess of the acid [hydric nitrate] employed."—*Miller*.

CHAPTER VII.

CHLORINE.

[*See page 121, Part I.*]

Symbol Cl = 35.5 ; molecular weight Cl_2 = 71 = 2 volumes.

Chlorine is usually set free from hydric chloride by manganic binoxide, heat being applied.

$$4(HCl) + MnO_2 = MnCl_2 + 2(H_2O) + Cl_2.$$

Sometimes, however, fused sodic chloride is used, and then hydric sulphate is employed to form hydric chloride from sodic chloride, and at the moment of its formation the hydrogen is taken away from it and chlorine is liberated.

$$MnO_2 + 2(NaCl) + 3(H_2SO_4) = 2(NaHSO_4) + MnSO_4 + 2(H_2O) + Cl_2.$$

In the first reaction only half the chlorine is evolved from hydric chloride. In the second, all the chlorine is evolved from the sodic chloride.

When chlorine acts as a disinfecting agent and decomposes hydric sulphate, sulphur is precipitated. This reaction will be given when treating of hydric sulphide.

If chlorine gas and steam be passed through a tube heated to a high temperature the following reaction occurs—

$$2(Cl_2) + 2(H_2O) = 4(HCl) + O_2.$$

HYPOCHLOROUS ACID.

[*See page 127, Part I.*]

Symbol Cl_2O; molecular weight 87, $Cl_2 = 71$, $O = 16$, $Cl_2O = 2$ volumes; density 43.5.

Chlorine gas passed over mercuric oxide gives the following reaction—

$$2Cl_2 + \underset{\text{(Mercuric oxide.)}}{HgO} = \underset{\text{(Mercuric chloride.)}}{HgCl_2} + Cl_2O.$$

HYDRIC HYPOCHLORITE.

[*See page 128, Part I.*]

Symbol $HClO$; molecular weight 52.5, $H = 1$, $Cl = 35.5$, $O = 16$.

When hypochlorous acid unites with water, hydric hypochlorite is formed, thus—

$$Cl_2O + H_2O = \underset{\text{(Hydric hypochlorite.)}}{Cl_2H_2O_2}.$$

$Cl_2H_2O_2$ is not, however, the molecule; but, as in hydric nitrate, half that quantity, hydric hypochlorite being monobasic—

$$\frac{Cl_2H_2O_2}{2} = HClO.$$

When chlorine is passed into water, hydric chloride and hydric hypochlorite are formed—

$$Cl_2 + H_2O = HCl + HClO.$$

The reaction which takes place when bleaching powder, calcic hypochlorite, is made, is as follows:—

$$\underset{\text{(Calcic hydrate.)}}{2(CaH_2O_2)} + 2Cl = \underset{\text{(Calcic chloride.)}}{CaCl_2} + 2(H_2O) + \underset{\text{(Calcic hypochlorite.)}}{CaCl_2O_2}.$$

When hydric chloride acts upon a strong solution of hydric hypochlorite, chlorine is liberated and water is formed, thus—

$$HCl + HClO = Cl_2 + H_2O.$$

HYDRIC CHLORATE.

[*See page 131, Part I.*]

Symbol $HClO_3$; molecular weight 84.5, H = 1, Cl = 35.5, O_3 = 48.

Chlorine gas, passed into a solution of potassic hydrate in the cold, reacts thus—

$$2(KHO) + Cl_2 = KCl + KClO + H_2O.$$

But if the solution of potassic hydrate be hot, then—

$$6(KHO) + 3Cl_2 = \underset{\text{(Potassic chloride.)}}{5(KCl)} + \underset{\text{(Potassic chlorate.)}}{KClO_3} + 3(H_2O).$$

If the solution formed in the cold be afterwards boiled, the equation representing the change which takes place is as follows:—

$$3(KClO) = 2(KCl) + KClO_3.$$

The water and potassic chloride are here omitted, because they take no part in the reaction.

Hydric chlorate can be obtained by acting on potassic chlorate with hydric fluosilicate—

$$\underset{\text{(Hydric fluosilicate.)}}{SiF_6H_2} + 2(KClO_3) = \underset{\text{(Potassic fluosilicate.)}}{SiF_6K_2} + 2(HClO_3).$$

Potassic fluosilicate is thrown down as a precipitate, while suspended in water it is opalescent, and can at first be hardly distinguished. When dry it is a brilliant white powder. Again, when hydric sulphate is added to a solution of barytic chlorate, insoluble barytic

sulphate is thrown down, and hydric chlorate remains in solution—

$$\underset{\text{(Barytic chlorate.)}}{Ba(ClO_3)_2} + H_2SO_4 = \underset{\text{(Barytic sulphate.)}}{BaSO_4} + 2(HClO_3).$$

Hydric chlorate is therefore monobasic. Barium, which is divalent, requires 2 molecules of it to react upon.

CHLOROUS ACID.

[*See page 134, Part I.*]

Symbol Cl_2O_3 ; molecular weight 119, $Cl_2 = 71$, $O_3 = 48$.

The formula for chlorous acid is Cl_2O_3. That of hydric chlorite is $HClO_2$.

$$Cl_2O_3 + H_2O = Cl_2H_2O_4, \text{ or } 2(HClO_2).$$

Other chlorites are formed by the replacement of hydrogen by metals, as $KClO_2$.

CHLORIC OXIDE.

[*See page 134, Part I.*]

The symbol for this oxide is Cl_2O_4.

HYDRIC PERCHLORATE.

[*See page 135, Part I.*]

Symbol $HClO_4$; molecular weight 100.5, $H = 1$, $Cl = 35.5$ $O_4 = 64$.

Hydric perchlorate can be obtained by distilling potassic perchlorate with dilute hydric sulphate. Potassic perchlorate is formed by the action of heat on potassic chlorate, thus—

$$2(KClO_3) = KCl + KClO_4 + O_2.$$

It is supposed that potassic chlorite, $KClO_2$, is first formed, and that this breaks up into KCl and O_2.

HYDRIC CHLORIDE.

[*See page 138, Part I.*]

Symbol HCl; molecular weight 36.5, H = 1, Cl = 35.5, HCl = 2 volumes; density 18.25.

Hydric chloride is prepared by the action of hydric sulphate on sodic chloride. The reaction at a moderate temperature is—

$$\underset{\text{(Sodic chloride.)}}{NaCl} + H_2SO_4 = \underset{\text{(Hydrosodic sulphate.)}}{NaHSO_4} + HCl.$$

But at a high temperature—

$$NaCl + NaHSO_4 = Na_2SO_4 + HCl.$$

It will be here seen that these reactions are similar to those which take place in making hydric nitrate.

Hydric chloride is monobasic; there is only one class of chlorides, just as there is only one class of nitrates or chlorates.

Tests for Chlorides.

When argentic nitrate is added to a soluble chloride, a precipitate of argentic chloride is formed, thus—

$$NaCl + \underset{\text{(Argentic nitrate.)}}{AgNO_3} = \underset{\text{(Sodic nitrate.)}}{NaNO_3} + AgCl.$$

or

$$CaCl_2 + 2(AgNO_3) = Ca(NO_3)_2 + 2(AgCl).$$

In the second equation, divalent calcium requires 2 molecules of argentic nitrate for the reaction.

When potassic dichromate and sodic chloride, or other soluble chloride and strong hydric sulphate are heated together, chlorochromic acid is formed, thus—

$$\underset{\text{Potassic dichromate.}}{K_2Cr_2O_7} + 4(NaCl) + 3(H_2SO_4) = K_2SO_4 + 2(Na_2SO_4) + \underset{\text{Chlorochromic acid.}}{2(CrCl_2O_2)} + 3(H_2O).$$

And when chlorochromic acid is passed into water—

$$\underset{\text{Chlorochromic acid.}}{CrCl_2O_2} + 2(H_2O) = \underset{\text{Chromic acid.}}{H_2CrO_4} + 2\,HCl.$$

The persistent yellow colour is due to the chromate.

Another test for a chloride, not noticed in Part I., consists in the addition of mercurous nitrate to a solution of a chloride.

$$2(NaCl) + \underset{\text{Mercurous nitrate.}}{Hg_2(NO_3)_2} = 2(NaNO_3) + \underset{\text{Mercurous chloride.}}{Hg_2Cl_2}.$$

The precipitate formed, mercurous chloride or calomel, is white, but, on the addition of ammonia, is blackened.

BROMINE.

[*See page 146, Part I.*]

Symbol Br $=80$; molecular weight $Br_2 = 160 = 2$ volumes.

Bromine is extracted from bromide of sodium, obtained from the mother liquid of seawater by the method described in Part I.; bromine is obtained free by a process similar to that employed for getting free chlorine. The equation is—

$$2(NaBr) + 2(H_2SO_4) + MnO_2 = MnSO_4 + Na_2SO_4 + 2(H_2O) + Br_2.$$

HYDRIC HYPOBROMITE.

[*See page 148, Part I.*]

Symbol $HBrO$; molecular weight 97, $H = 1$, $Br = 80$, $O = 16$.

The action of bromine on mercuric oxide in the presence of water is thus expressed—

$$HgO + 2\,Br_2 + H_2O = \underset{\text{(Mercuric bromide.)}}{HgBr_2} + 2(HBrO).$$

HYDRIC BROMATE.

[*See page 149, Part I.*]

Symbol $HBrO_3$; molecular weight 129, $H = 1$, $Br = 80$, $O_3 = 48$.

The action of bromine on potassic hydrate is the same as that of chlorine, and is represented by the same equations, with the change of Cl for Br. When chlorine gas is passed through bromine water, the reaction which occurs may be represented as follows:—

$$Br_2 + 5\,Cl_2 + 6(H_2O) = 10(HCl) + 2\underset{\text{(Hydric bromate.)}}{(HBrO_3)}.$$

HYDRIC BROMIDE.

[*See page 149, Part I.*]

Symbol HBr; molecular weight 81, $H = 1$, $Br. = 80$, $HBr = 2$ volumes; density 40.5.

If water be made to act on phosphoric bromide, hydric bromide and hydric phosphate are formed, thus—

$$\underset{\text{(Phosphoric bromide.)}}{PBr_5} + 4(H_2O) = \underset{\text{(Hydric phosphate.)}}{H_3PO_4} + 5\underset{\text{(Hydric bromide.)}}{(HBr.)}$$

HYDRIC IODATE.

[*See page 155, Part I.*]

Symbol HIO_3; molecular weight 176, $H = 1, I = 127, O_3 = 48$.

The action of chlorine on iodine suspended in water is thus represented—

$$I_2 + 5Cl_2 + 6\,(H_2O) = 10\,(HCl) + \underset{\text{(Hydric iodate.)}}{2(HIO_3)}$$

POTASSIC IODATE.

[*See page 155, Part I.*]

When iodine is boiled with potassic hydrate, potassic iodide, potassic iodate, and water are formed.

$$3I_2 + 6(KHO) = \underset{\text{(Potassic iodide.)}}{5(KI)} + \underset{\text{(Potassic iodate.)}}{KIO_3} + 3(H_2O).$$

If chlorine gas be passed into a solution of potassic hydrate containing iodine, potassic chloride and potassic iodate are produced together with water.

$$I_2 + 12(KHO) + 5Cl_2 = 10(KCl) + 2(KIO_3) + 6(H_2O).$$

HYDRIC IODIDE.

[*See page 157, Part I.*]

Symbol HI; molecular weight 128, $H. = 1, I = 127$, $HI = 2$ volumes; density 64.

When hydric sulphide is passed through water in which iodine is suspended, it is decomposed, hydric iodide being formed and sulphur precipitated—

$$I_2 + \underset{\text{(Hydric sulphide.)}}{H_2S} = \underset{\text{(Hydric iodide.)}}{2(HI)} + \overset{*}{S}.$$

The water is not noticed in the equation, as it takes no part in the reaction.

Phosphorous iodide, acted upon by water, is decomposed—hydric iodide and hydric phosphite being formed—

$$\underset{\left(\begin{smallmatrix}\text{Phosphorous}\\ \text{iodide.}\end{smallmatrix}\right)}{PI_3} + 3(H_2O) = H_3PO_3 + 3(HI).$$

POTASSIC IODIDE.

[*See page 158, Part I.*]

The equation representing the action of iodine on potassic hydrate has been given under potassic iodate. When potassic iodide and iodate are heated together the oxygen is driven off from the iodate—

$$10\,(KI) + 2\,(KIO_3) = 12(KI) + 3\,O_2.$$

FLUORINE.

[*See page 162, Part I.*]

Symbol F = 19.

Calcic fluoride heated with hydric sulphate gives calcic sulphate and hydric fluoride thus—

$$\underset{\left(\begin{smallmatrix}\text{Calcic}\\ \text{fluoride.}\end{smallmatrix}\right)}{CaF_2} + H_2SO_4 = CaSO_4 + \underset{\left(\begin{smallmatrix}\text{Hydric}\\ \text{fluoride.}\end{smallmatrix}\right)}{2(HF)}.$$

CHAPTER VIII.

SULPHUROUS ACID.

[*See page 169, Part I.*]

Symbol SO_2; molecular weight 64, $S = 32$, $O_2 = 32$, $SO_2 = 2$ volumes; density 32.

When hydric sulphate is heated with mercury the following decomposition takes place—

$$\overset{*}{Hg} + 2(H_2SO_4) = \underset{\text{(Mercuric sulphate.)}}{HgSO_4} + 2(H_2O) + \underset{\text{(Sulphurous acid.)}}{SO_2}.$$

The action of carbon on hydric sulphate, aided by heat, may be expressed thus—

$$\overset{*}{C} + 2(H_2SO_4) = CO_2 + 2(H_2O) + 2(SO_2).$$

Sulphurous acid combines with bases, forming two classes of salts, for example—

(Hydro-potassic sulphite.)	(Potassic sulphite.)
$KHSO_3$	K_2SO_3

Sulphurous acid dissolved in water may be regarded as—

$$H_2SO_3 \text{ and water}$$
$$\text{for } H_2O + SO_2 = H_2SO_3,$$

and both these atoms of hydrogen can be replaced by metals—

$$\begin{array}{l} K|H \\ \qquad SO \\ \quad H \end{array}$$

$$\begin{array}{l} K|H \\ \qquad SO_3 \\ K|H \end{array}$$

SULPHURIC ACID.

[*See page 171, Part I.*]

Symbol SO_3; molecular weight 80, $S = 32$, $O_3 = 48$, $SO_3 = 2$ volumes; density 40.

Nordhausen acid has the formula $H_2S_2O_7$, that is—

$$\underset{H_2SO_4}{\left(\begin{array}{c}\text{Hydric}\\ \text{sulphate.}\end{array}\right)} \quad \text{and} \quad \underset{SO_3.}{\left(\begin{array}{c}\text{Sulphuric}\\ \text{acid.}\end{array}\right)}$$

The action of heat is to drive off SO_3, which is collected in a cooled receiver.

If hydrosodic sulphate be heated in a glass retort, the following change takes place :—

$$2(NaHSO_4) = \underset{\left(\begin{array}{c}\text{Sodic}\\ \text{disulphate.}\end{array}\right)}{Na_2S_2O_7} + H_2O \text{ (which is expelled).}$$

If this sodic disulphate be heated to a high temperature in a retort made of fireclay, it breaks up thus—

$$Na_2S_2O_7 = Na_2SO_4 + SO_3.$$

Hydric sulphate is represented by the symbol H_2SO_4 it is made by the union of water and sulphuric acid.

$$H_2O + SO_3.$$

When phosphoric acid acts upon this, it takes away H_2O, and SO_3 is set free.

HYDRIC SULPHATE.

[*See page 175, Part I.*]

Symbol H_2SO_4; molecular weight 98, $H_2 = 2$, $S = 32$, $O_4 = 64$.

In the manufacture of hydric sulphate, sulphurous acid is first formed; this meets with nitrous acid, takes some of its oxygen, and becomes sulphuric acid, and this unites with water, forming hydric sulphate. The first action is represented thus—

$$SO_2 + N_2O_3 = 2(NO) + SO_3,$$

and the second—

$$SO_3 + H_2O = H_2SO_4.$$

Hydric sulphate of specific gravity 1.842 has the composition H_2SO_4. But if water be added until its specific gravity reaches 1.78, a hydrate is formed which crystallises between 7°C. and 9°C. Its composition answers to the formula H_2SO_4, H_2O. Professor Graham has obtained another hydrate by evaporating a dilute solution of hydric sulphate at 100°C.; the density of this body is found to be 1.63, and its formula is—

$$H_2SO_4,\ 2(H_2O).$$

The hydrogen in hydric sulphate can be replaced wholly or in part by a monobasic metal, thus we have—

$$\left.\begin{array}{r}\text{Hydro-potassic sulphate } KHSO_4\\ \text{Potassic sulphate } K_2SO_4\end{array}\right\}\begin{array}{l}\text{Potassium being}\\ \text{monovalent}\end{array}$$

Calcic sulphate $CaSO_4$ calcium being divalent

Mercuric sulphate $HgSO_4$ mercury being divalent

Hydric sulphate is therefore dibasic.

Sulphates can be decomposed by heating them with charcoal (common coal-dust answers well); they are reduced to sulphides, and these, when acted upon by hydric chloride, give off hydric sulphide.

$$BaSO_4 + 4\overset{*}{C} = BaS + 4CO$$
$$\text{and } BaS + 2HCl = H_2S + BaCl_2.$$

HYDRIC HYPOSULPHITE.

Hyposulphurous acid, S_2O_2, is not known in the free state.

Hydric hyposulphite, whose molecule is $H_2S_2O_3$, is very unstable. It decomposes so easily that sulphur is deposited and sulphurous acid is set free, when sodic hyposulphite is acted upon by hydric chloride. Sodic hyposulphite is used by photographers for dissolving undecomposed argentic iodide (*i.e.* iodide of silver which has not been acted upon by light); in fact this hyposulphite has the property of dissolving salts of silver which have not been acted upon by light. Sodic hyposulphite is made by passing sulphurous acid gas through a solution of sodic sulphide. Rose says that the hyposulphites contain the elements of water in addition to the elements given in the above formula, therefore the soda salt is $Na_2H_2S_2O_4$.

HYDRIC SULPHIDE.

[*See page 180, Part I.*]

Symbol H_2S; molecular weight 34, $H_2 = 2$, $S = 32$, $H_2S = 2$ volumes; density 17.

Hydric sulphide is formed by acting upon ferrous sulphide with dilute hydric sulphate—

$$\underset{\left(\begin{smallmatrix}\text{Ferrous}\\\text{sulphide.}\end{smallmatrix}\right)}{FeS} + H_2SO_4 = FeSO_4 + \underset{\left(\begin{smallmatrix}\text{Hydric}\\\text{sulphide.}\end{smallmatrix}\right)}{H_2S}.$$

Or, by heating antimonious sulphide with hydric chloride—

$$\underset{\left(\begin{smallmatrix}\text{Antimonious}\\\text{sulphide.}\end{smallmatrix}\right)}{Sb_2S_3} + 6(HCl) = \underset{\left(\begin{smallmatrix}\text{Antimonious}\\\text{chloride.}\end{smallmatrix}\right)}{2(SbCl_3)} + 3(H_2S).$$

Antimony is a trivalent element, therefore the molecule of hydric chloride, HCl, must be taken six times for the decomposition of the molecule of antimonious sulphide Sb_2S_3. Moreover, the fact that the molecule of antimonious sulphide contains 3 atoms of sulphur, each of which is divalent, shows that enough hydric chloride must be taken to furnish 6 atoms of hydrogen to combine with them.

When hydrogen gas is evolved by the usual method, if hydric sulphate be added undiluted, a small quantity of sulphurous acid is formed, thus—

$$\overset{*}{Zn} + 2(H_2SO_4) = ZnSO_4 + 2(H_2O) + SO_2,$$

and the hydrogen evolved on the addition of water decomposes the sulphurous acid, forming hydric sulphide, thus—

$$SO_2 + 3H_2 = 2(H_2O) + H_2S.$$

When passed into chlorine water, hydric sulphide is decomposed with precipitation of sulphur—

$$Cl_2 + H_2S = 2(HCl) + \overset{*}{S}.$$

Hydric sulphide is used in the laboratory to precipitate metals, as sulphides, from solutions of their salts.

The molecule of cupric sulphate is $CuSO_4$, or $CuOSO_3$. If cupric sulphate in solution be acted upon by hydric sulphide, the sulphur of the sulphide replaces the oxygen of the cupric oxide, and the hydrogen (2 atoms) takes the place of divalent copper. Cupric sulphide and hydric sulphate being formed.

$$CuSO_4 + H_2S = CuS + H_2SO_4,$$

or

$$CuOSO_3 + H_2S = CuS + H_2OSO_3.$$

Compound of Sulphur and Carbon.

CARBONIC SULPHIDE.

[*See page 184, Part I.*]

Symbol CS_2; molecular weight 76, $C=12$, $S_2=64$, $CS_2=2$ volumes; density 38.

When sulphur vapour is passed over red-hot charcoal, carbonic sulphide is formed. When oxygen is passed over red-hot charcoal, carbonic acid is formed. The similarity between sulphur and oxygen is here very apparent—

$$C + S_2 = CS_2, \quad C + O_2 = CO_2.$$

The density of oxygen is 16; its compound with carbon is a gas. The density of sulphur vapour is 32; its compound with carbon is a liquid. Sulpho-carbonic acid (carbonic sulphide) combines with sulphur bases (sulphides) forming sulpho-carbonates, just as carbonic acid unites with oxygen bases (oxides), forming carbonates. The molecule of potassic carbonate is K_2CO_3. The molecule of potassic sulpho-carbonate is

K_2CS_3, *i.e.* K_2S,CS_2, just as the carbonate is K_2O,CO_2. When these sulpho-carbonates are decomposed by hydric chloride, in presence of water, hydric sulpho-carbonate is formed, thus—

$$K_2CS_3 + 2(HCl) = 2(KCl) + \underset{\text{(Hydric sulpho-carbonate.)}}{H_2CS_3}.$$

CHAPTER IX.

PHOSPHORUS.

[*See page 184, Part I.*]

Symbol $P = 31$; molecular weight $P_4 = 124, = 4$ volumes; density 62.

The reaction which takes place when bone earth is treated with dilute hydric sulphate is—

(Tricalcic phosphate or bone earth.) (Calcic sulphate.) (Calcic superphosphate.)

$$Ca_3P_2O_8 + 2(H_2SO_4) = 2(CaSO_4) + CaH_4P_2O_8.$$

And when calcic superphosphate is heated with charcoal phosphorus distils over; the reaction may be thus expressed—

$$3(CaH_4P_2O_8) + C_{10} = Ca_3P_2O_8 + 6(H_2O) + 10(CO) + P_4.$$

If the temperature be raised very high, some water may be decomposed and hydrogen set free.

Phosphoric Acid and Trihydric Phosphate.

PHOSPHORIC ACID.

[*See page 187, Part I.*]

Symbol P_2O_5; molecular weight 142, $P_2 = 62$, $O_5 = 80$.

When phosphorus is burnt in excess of oxygen, phosphoric acid is formed. Its formula is P_2O_5; and when this is acted upon by water

$$P_2O_5 + 3(H_2O) = \underset{\left(\substack{\text{Trihydric}\\\text{phosphate.}}\right)}{2(H_3PO_4)}.$$

When trihydric phosphate is made from calcic superphosphate by the action of ammonia—

$$3(CaH_4P_2O_8) + 12(NH_3) = Ca_3P_2O_8 + \underset{\left(\substack{\text{Triammonic}\\\text{phosphate.}}\right)}{4\{(NH_4)_3PO_4\}}$$

And the triammonic phosphate, when heated, loses its ammonia, thus—

$$(NH_4)_3PO_4 = H_3PO_4 + 3(NH_3).$$

And on further heating, trihydric phosphate becomes what is called glacial phosphoric acid, or hydric metaphosphate—

$$H_3PO_4 \quad = \quad \underset{\left(\substack{\text{Metaphos-}\\\text{phate.}}\right)}{HPO_3} + H_2O.$$

And on boiling the metaphosphate with water for some time, the reverse action takes place.

The molecule of dihydro sodic phosphate is NaH_2PO_4

ditto	Hydro-disodic phosphate or rhombic phosphate of soda	Na_2HPO_4
ditto	Trisodic phosphate	Na_3PO_4

Microcosmic salt is represented thus, $NaHNH_4PO_4$, and ammonio magnesic phosphate thus, $MgNH_4PO_4$. Here magnesium being divalent, its atom takes the place of 2 atoms of hydrogen.

When argentic nitrate (neutral solution) is added to a solution of hydro-disodic phosphate which is also neutral, a yellow precipitate is formed, but the liquid is acid to test paper; this is owing to hydric nitrate being set free in the reaction which takes place.

$$3(AgNO_3) + Na_2HPO_4 = \underset{\text{(Triargentic phosphate.)}}{Ag_3PO_4} + 2(NaNO_3) + HNO_3$$

TETRAHYDRIC DIPHOSPHATE

Is also called pyrophosphoric acid. This latter name is given because it is formed by the agency of heat. When rhombic phosphate of soda is heated to a bright red heat—

$$2(Na_2HPO_4) = \underset{\text{(Tetrasodic diphosphate.)}}{Na_4P_2O_7} + H_2O.$$

If this tetrasodic diphosphate be dissolved in water it crystallises in needle-shaped crystals. It gives a *white* precipitate with argentic nitrate, no hydric nitrate being set free, as is the case when argentic nitrate is added to hydro-disodic phosphate.

$$4(AgNO_3) + Na_4P_2O_7 = 4(NaNO_3) + Ag_4P_2O_7.$$

If a solution of plumbic acetate be added to a solution of tetrasodic diphosphate, a white precipitate of diplumbic diphosphate, $Pb_2P_2O_7$, is formed; lead, being divalent, replaces 2 atoms of sodium by 1 atom. This white precipitate being insoluble, can only be mechanically mixed with water, but if through such a mechanical mixture hydric sulphide be passed, the lead will be precipitated as sulphide and tetrahydric diphosphate will be formed and remain in the solution. The action which takes place may be thus represented—

$$Pb_2P_2O_7 + 2(H_2S) = \underset{\text{(Plumbic sulphide.)}}{2(PbS)} + \underset{\text{(Tetrahydric diphosphate.)}}{H_4P_2O_7}.$$

This tetrahydric diphosphate, if boiled with water, is decomposed thus—

$$H_4P_2O_7 + H_2O = 2(H_3PO_4).$$

To go more fully into this interesting subject would take us beyond our present limits. It is hoped that the student will observe attentively the different actions of heat on phosphates in their different combinations.

HYDRIC METAPHOSHATE.

When microcosmic salt is heated, sodic metaphosphate is formed—

$$NaHNH_4PO_4 = \underset{\text{(Sodic metaphosphate.)}}{NaPO_3} + NH_3 + H_2O,$$

also, when NaH_2PO_4 is ignited, it breaks up, thus—

$$NaH_2PO_4 = NaPO_3 + H_2O.$$

The hydrogen salt is obtained by the action of plumbic acetate and hydric sulphide on sodic metaphosphate, as explained in the case of tetrahydric diphosphate. The true formula for this salt is not known. Whether its molecule is HPO_3, or some multiple of this, is doubtful. Other particulars regarding it have been already given.

Phosphorous Acid and Hydric Phosphite.

PHOSPHOROUS ACID.

[*See page 193, Part I.*]

Symbol P_2O_3; molecular weight 110, $P_2 = 62$, $O_3 = 48$.

Phosphorous acid when acted upon by water gives hydric phosphite, thus—

$$P_2O_3 + 3(H_2O) = \underset{\text{(Hydric phosphite.)}}{2(H_2HPO_3)}.$$

When phosphorous chloride acts upon water the decomposition is represented thus—

$$PCl_3 + 3(H_2O) = 3(HCl) + H_2HPO_3.$$

Hydric phosphite, contains 3 atoms of hydrogen; only two of these can be replaced by potassium or a monovalent element, thus—hydropotassic phosphite is KH_2PO_3; potassic phosphite, which is neutral, is K_2HPO_3. Hydric phosphite, therefore, is dibasic. It acts as a reducing agent—thus, with mercuric chloride the following reaction takes place—

$$2(HgCl_2) + H_2HPO_3 + H_2O = Hg_2Cl_2 + H_3PO_4 + 2(HCl).$$

Professor Miller gives a characteristic reaction in the reduction of hydric sulphite—

$$3(H_2HPO_3) + H_2SO_3 = 3(H_3PO_4) + H_2S.$$

If hydric phosphite be heated, it is decomposed into hydric phosphide and trihydric phosphate.

HYDRIC PHOSPHIDE.

[*See page 194, Part I.*]

Symbol PH_3; molecular weight 34, $P=31$, $H_3=3$, PH_3 = 2 volumes; density 17.

When hydric phosphite is heated—

$$4(H_2HPO_3) = \underset{\text{(Hydric phosphide.)}}{PH_3} + 3(H_3PO_4).$$

Hydric phosphide is also made by heating phosphorus in a solution of sodic hydrate—

$$P_4 + 3(NaHO) + 3(H_2O) = \underset{\text{(Sodic hypophosphite.)}}{3(NaH_2PO_2)} + PH_3.$$

The gas made in this manner ignites immediately it comes in contact with air or oxygen; this is owing to the presence of another hydrogen compound of phosphorus, to which is attributed the formula P_2H_4, and which, when condensed, is a volatile liquid.

Compounds of Phosphorus and Chlorine.

PHOSPHOROUS CHLORIDE.

Symbol PCl_3; molecular weight 137.5, $P = 31$, $Cl_3 = 106.5$, $PCl_3 = 2$ volumes; density 68.5.

Phosphorous chloride is a liquid. It fumes in air, dissolves phosphorus, and is decomposed by water. It is prepared by passing dry chlorine gas through melted phosphorus. The operation should be conducted in a retort, and if heat be applied the phosphorous chloride distils over as long as phosphorus is in excess, and is condensed in a receiver.

PHOSPHORIC CHLORIDE.

Symbol PCl_5; molecular weight 208.5, $P = 31$, $Cl_5 = 177.5$.

If phosphorus be treated with excess of chlorine, phosphoric chloride is formed. Phosphoric chloride is a yellowish-white crystalline powder. It is decomposed by heat into chlorine and phosphorous chloride. Water acts upon it in two ways. With a small quantity of water the reaction is as follows—

$$PCl_5 + H_2O = \underset{\text{(Phosphoric oxychloride.)}}{POCl_3} + 2(HCl.)$$

The body POCl is called chloro-phosphoric acid, or

oxychloride of phosphorus. With a large quantity of water the reaction is

$$PCl_5 + 4(H_2O) = H_3PO_4 + 5HCl.$$

Professor Williamson considers that this change takes place in successive stages. Thus, the formation of chloro-phosphoric acid takes place first, and this reacts on another molecule of water, thus—

$$POCl_3 + H_2O = PO_2Cl + 2(HCl);$$

and PO_2Cl again reacts on another molecule of water, thus—

$$PO_2Cl + H_2O = HPO_3 + HCl.$$

The hydric metaphosphate immediately taking up a molecule of water, $HPO_3 + H_2O$ becomes H_3PO_4.

CHAPTER X.

BORACIC ACID.

[*See page 196, Part I.*]

Symbol B_2O_3; molecular weight 70, $B_2 = 22$, $O_3 = 48$.

Boron forms one compound with oxygen.

Hydric borate has the formula H_3BO_3, it is obtained from borax, $Na_2B_4O_7$, by acting upon it with hydric sulphate.

Boric chloride, BCl_3, is formed by mixing boracic acid and charcoal, and heating them in a hard-glass tube whilst a current of chlorine gas is passing through. It is a clear liquid, easily decomposed by water.

SILICIC ACID.

[*See page 200, Part I.*]

Symbol SiO_2; molecular weight 60, $Si = 28$, $O_2 = 32$.

Silicic acid forms many different silicates, which are too complicated for consideration here. Two, however, may be noticed—dihydric silicate, H_2SiO_3; and tetrahydric silicate, H_4SiO_4. Dihydric silicate is obtained when an aqueous solution of it is evaporated in vacuo at about 26°C. It should be dried over hydric

sulphate for a few days after evaporation; it then has nearly the composition H_2SiO_3.—*Miller.* The gelatinous precipitate formed when hydric chloride is added to a solution of potassic or sodic silicate, is supposed to have the composition H_4SiO_4. Some natural silicates are found to have the same composition—for instance, the hydrogen in H_4SiO_4 is replaced by magnesium in the well-known silicate meerschaum, Mg_2SiO_4 (magnesium being divalent, its atom replaces 2 atoms of hydrogen). Silicate of lime also has the formula Ca_2SiO_4. Another silicate of lime, called Wollastonite, has the formula $CaSiO_3$. This is a silicate of the same composition as dihydric silicate.

SILICIC CHLORIDE.

[*See page 202, Part I.*]

Symbol $SiCl_4$; molecular weight 170, $Si = 28$, $Cl_4 = 142$, $SiCl_4 = 2$ volumes; density 85.

In this compound 4 atoms of chlorine take the place of 2 atoms of oxygen in silicic acid. When chlorine is passed over silica and carbon, as described in Part I., this reaction takes place—

$$SiO_2 + 2Cl_2 + 2C = SiCl_4 + 2(CO).$$

When silicic chloride comes in contact with water, it is immediately decomposed, thus—

$$SiCl_4 + 2(H_2O) = SiO_2 + 4(HCl).*$$

* The silicic acid combines with water, forming a silicate.

SILICIC FLUORIDE.

Symbol SiF_4; molecular weight 104, $Si = 28$, $F_4 = 76$, $SiF_4 = 2$ volumes; density 52.

Silicic fluoride is a gas without colour, which fumes strongly when it comes in contact with air; it is formed when hydric fluoride acts on silica or glass.

$$\underset{\left(\begin{array}{c}\text{Hydric}\\\text{fluoride.}\end{array}\right)}{4(HF)} + SiO_2 = SiF_4 + 2(H_2O).$$

The best way to prepare silicic fluoride is by mixing together calcic fluoride, powdered glass, and strong hydric sulphate. The action commences without heat; but after a time a very gentle heat should be applied. In making this gas a common Florence flask may be used, if excess of powdered glass be employed, as the hydric fluoride first formed by the action of hydric sulphate on calcic fluoride attacks the powdered glass before the glass of the flask. However, in time that often becomes destroyed. A platinum or lead vessel is better fitted for the purpose. The decomposition which takes place is as follows:—

$$2(CaF_2) + 2(H_2SO_4) + SiO_2 = 2(CaSO_4) + 2(H_2O) + SiF_4.$$

This gas must be collected over mercury. Silicic fluoride was condensed to a liquid by Faraday. When it is passed into water a decomposition takes place. One-third of the silicon unites with the oxygen of the water and is thrown down in white flakes, the remaining silicon and fluorine take into combination

the hydrogen of the water decomposed by the silicon, thus—

$$\underset{}{3(SiF_4) + 2(H_2O) = 2\overset{\left(\substack{\text{Hydric}\\\text{fluosilicate.}}\right)}{(H_2SiF_6)} + SiO_2.}$$

The substance formed is called hydric fluosilicate; its solution is an acid liquid, its hydrogen can be replaced by metals. K_2SiF_6 is the molecule of potassic fluosilicate; $BaSiF_6$ is that of barytic fluosilicate. Potassic fluosilicate is the substance employed in obtaining silicon, as described in the formation of crystallised silicon.

CHAPTER XI.

Some points with respect to "valency" could not be noticed before, as the bodies which illustrate them had not been treated of. In ammonia, NH_3, nitrogen is trivalent, but in ammonic chloride, NH_4Cl, the nitrogen is in combination with 4 atoms of hydrogen and 1 of chlorine; it is therefore pentavalent, *i.e.* FIVE valent. In hydric phosphide, PH_3, phosphorus is trivalent, but in phosphoric pentachloride, PCl_5, it is pentavalent, There are also instances of this double valency of an element in other elements of uneven valency, but as they do not come under our notice in this work, it is needless to refer to them.

Oxygen in carbonic acid, CO_2, is divalent, but in carbonic oxide, CO, it is tetravalent—or carbon is divalent; the latter seems likely to be the case. Oxygen is almost invariably divalent. Silicon is tetravalent, as is seen in SiO_2.

It must be borne in mind, that the explanation of properties by such terms as "valency" is only a convenient method of stating what we know about them, and that the use of such terms may be only of short duration. It may be discovered that some of those substances which we now regard as elements, are not

elements at all, but compounds. And even advanced knowledge may substitute some theory which may have a more sure foundation than the atomic theory. It is well to remark this, as young minds are often apt to espouse warmly views of a certain school, and to regard as facts what are nothing more than assumptions. It is generally believed that the division of elements into those of even and uneven valency is one that will stand; however, the body nitric oxide, NO, is one which, for the present at least, throws a doubt on this opinion, for nitrogen is usually trivalent and oxygen divalent, and *here* one or other of them must change its valency. It will be remembered that NO occupies 2 volumes, and this is strong evidence that that quantity of nitric oxide is the molecule.

Molecules are represented symbolically by the formulæ which truly express their composition, and as symbols represent definite weights of an element, the elements must be expressed by their symbols as many times as is necessary to show what weight of them is in the molecule of any compound. The molecule of hydric nitrate contains only 1 atom of hydrogen, therefore there is but 1 H in its formula, HNO_3. That of hydric sulphate contains 2 atoms of hydrogen, therefore there are 2Hs in its formula H_2SO_4. In the molecule of phosphorus there are 4 atoms, so that P_4 is the symbol of its molecule, and P_4 occupies 2 volumes. The atomic weight of phosphorus is 31, therefore $31 \times 4 = 124$ is the weight which occupies 2 volumes, and half this weight, viz. 62, is the density of phosphorus on the hydrogen scale. The molecule of sulphur is different according to the temperatures at which its

vapour density is taken. Thus, at 500°C. its molecule is S_6, that is, 6 atoms occupy 2 volumes. At a temperature of 1000°C. the molecule is S_2 or $S_2 = 2$ volumes, and its vapour density is therefore 32. An element which does not belong to the group we have undertaken to treat of here, but which has been alluded to several times, is another exception to the general rule that the molecule of an element consists of 2 atoms. The atom of mercury weighs 200, and this weight of it occupies 2 volumes, so that the atom and molecule are the same, and are represented by the symbol Hg; and the density of mercury vapour is half its atomic weight.

Acids, Bases, and Salts.—Silica is an acid, potassic oxide is a base, and potassic silicate is a salt. Under ordinary circumstances, silicic acid manifests no acid properties; it does not taste sour, it does not turn blue litmus paper red. Potassic oxide does not change the colour of red litmus paper; to put it into a condition to do so, changes it into another body—potassic hydrate. It is clear, then, that the manifestations of *certain* properties attributed to acids and bases are not distinctive marks of these bodies. That silica is an acid is clear, because it can displace sulphuric acid at high temperatures; for sodic sulphate is now largely used in the manufacture of glass, and the molecule SiO_2 displaces SO_3.

Sulphuric acid does not turn blue litmus paper red unless it be brought into contact with it, which is effected by mixing it with water; but in this case the molecule of water H_2O is a base, and it unites with a molecule of the acid, SO_3, forming the acid salt, whose molecule is H_2SO_4. If, instead of water, potassic hydrate unite

with sulphuric acid, a molecule of hydro-potassic sulphate $KHSO_4$ is formed, and this is an acid salt; but not so acid as hydric sulphate, because potassic hydrate is a more powerful base than water. Potassic oxide is a more powerful base still, and therefore potassic sulphate, whose molecule is K_2SO_4, exhibits neither acid nor basic (or alkaline) properties, for the base neutralises the acid. If sulphuric acid be made to react upon barytic oxide, a direct union takes place, and barytic sulphate, whose molecular formula is $BaSO_4$, is formed; this is a synthetical action. The formation of salts by double decomposition is such as takes place when the hydrogen of an acid salt is replaced by the more active metal of another salt; for example—

$$H_2SO_4 + 2(KHO) = K_2SO_4 + 2(H_2O)$$

or,

$$H_2SO_4 + 2(NaCl) = Na_2SO_4 + 2(HCl).$$

The common substantive use of the word "acid" makes it very difficult to give a precise definition of what is meant by the word "acid" in its chemical sense. An acid may be said to be a body which, in combination with a base, produces a salt, which salt may have either sour properties (*i.e.* it may taste sour or redden blue litmus paper), basic or alkaline properties, as in tri-hydric phosphate (*i.e.* it may have a soapy taste, and turn red litmus paper blue), or it may have neutral properties, and have a taste which is generally characterised by the word "saline."

A base is a body with opposite properties to that which an acid possesses, and which can, according to the strength of those properties, neutralise the properties

of an acid, when it is made to combine chemically with it.

A salt is the product of the direct union of an acid and base, or is formed by double decomposition when salts which can react are brought into contact with one another.

Very much more might be said on this subject, but in a work like the present, its further discussion would be unadvisable.

CHAPTER XII.

SPECIFIC WEIGHTS OF GASES.

(Air being taken as 1.)

Hydrogen	0.06921
Oxygen	1.1056
Sulphur vapour ...	2.112
Nitrogen	0.974
Steam	0.622
Carbonic acid ...	1.529
Carbonic oxide ...	0.967
Marsh gas	0.5528
Chlorine	2.453
Bromine (vapour)	5.528
Iodine (vapour) ...	8.756
Sulphurous acid ...	2.2112
Hydric sulphide ...	1.1674
Phosphorus (vapour)	4.284
Hydric chloride (gas)	1.261
Nitrous oxide... ...	1.528
Nitric oxide	1.039
Ammonia (gas) ...	0.5896
Olefiant gas... ...	.967

SPECIFIC WEIGHTS OF SOLIDS AND LIQUIDS.

(Water being taken as 1.)

Carbon—	
As diamond	3.33 to 3.55
As graphite	2.35 to 2.15
Bromine at 0°C. ...	3.187
Iodine (solid)	4.947
Sulphur—	
Octohedral	2.05
Prismatic	1.98
Plastic...	1.96
Hydric sulphate (oil of vitriol)	1.842
$H_2SO_4H_2O$	1.78
$H_2SO_42(H_2O)$ ...	1.62
Phosphorus (red) ...	2.14
Ditto (clear) ...	1.83
Hydric chloride (saturated solution) ...	1.2474
Hydric nitrate at 18°C.	1.517
Ammonia (solution), about	.8844
Boron	2.68
Silicon	2.49
Silicic acid, as quartz	2.642
Amorphous	2.3

Metrical Weights and Measures.

(Compiled and verified expressly for this work.)

The standard of French measures, commonly called the metrical system, is the mètre, which is the ten-millionth part of the distance from the equator to either pole, as ascertained by actual measurement of an arc of the meridian.

The mètre, the unit for measures of length = 39.37079 inches.

The are, the unit for measures of surfaces = 100 square mètres = 119.60332602 square yards.

The stère, the unit for measures of solids = a cubic mètre = 35.31658074 cubic feet = 1000 litres = 220.0966767 imperial gallons.

The litre, the unit for measures of capacity = a cubic décimètre, or cubic tenth part of a mètre = a millistère = 61.027051519 cubic inches = 1.760773414 pints.

The gramme, the unit for measures of weight = the weight of a cubic centimètre, or cubic hundredth part of a mètre of distilled water at 4°C. under normal pressure = 15 4323488 English grains.

In order to express the decimal proportion, the following prefixes to the principal units have been adopted :—

Déca signifies ten times the unit		
Hecto	,,	one hundred times the unit
Kilo	,,	one thousand ,, ,,
Myria	,,	ten thousand ,, ,,
Déci expresses the tenth part of the unit		
Centi	,,	hundredth part of the unit
Milli	,,	thousandth ,,

Measures of Length.

1 millimètre	= .001 mètres	= .03937079 inches
1 centimètre	= .01 ,,	= .3937079 ,,
1 inch	= .02539954 mèt.	= 1. ,,
1 décimètre	= .1 ,,	= 3.937079 ,,
1 foot	= .30479449 ,,	= 12. ,,
1 yard	= .91438348 ,,	= 36. ,,
1 mètre	= 1. ,,	= 39.37079 ,, or 3.2808992 feet
1 décamètre	= 10. ,,	= 10.936331 yards
1 hectomètre	= 100. ,,	= 109.36331 ,,
1 kilomètre	= 1000. ,,	= 1093.6331 ,,
1 mile	= 1609.3149261 ,,	= 1760. ,,
1 myriamètre	= 10000. ,,	= 6.2138242 miles

Measures of Surfaces.

1 square inch	= 6.4513669 square centimètres	
1 square foot	= 9.2899683 square decimètres	
	ares	
1 square yard	= .0083609715 ,,	= 9. square feet
1 centiare or 1 sq. mètre	= .01 ,,	= 10.76429934183 square feet
1 are or 100 square mètres	= 1. ,,	= 119.60332602 sq. yards
1 acre	= 40.467102 ,,	= 4840.
1 hectare	= 100. ,,	= 2.4711430996 acres
1 square mile	= 25898.94531 ,,	= 640

Measures of Capacity and Solids.

1 millilitre or cub. centimètre	= .001 litres	= .061027051519‍4 cubic inches
1 centilitre or 10 cb. centimètres	= .01 ,,	= .610270515194 cubic inches

1 cubic inch	= .0163861759 litres	= 1. cubic inch
1 décilitre or 100 cb. centimètres	= .1 ,,	= 6.1027051519 4 cubic inches
1 pint	= .56793224 ,,	= 34.65923 cubic inches
1 litre or millistère	= 1. ,,	= 61.0270515194 cubic inches or 1.760773437 pints
1 imperial gallon	= 4.543458 ,,	= 277.27384 cubic inches
1 décalitre or centistère	= 10. ,,	= .353165807 cubic feet
1 cubic foot	= 28.3153119 ,,	= 1.
1 hectolitre or decistère	= 100. ,,	= 3.353165807 cubic feet
1 cubic yard	= 764.513422 ,,	= 27. cubic feet
1 kilolitre, or 1 stère or 1 cubic mètre	= 1000. ,,	= 1.308021509 cubic yards

The imperial gallon is a measure containing 70,000 grains of distilled water at 62°F., but this measure of distilled water at 4°C., according to Mr. J. C. Brough's calculations, would weigh 70116.3636 grains.

Measures of Weight.

1 milligramme	= .001 grammes	= .015432348 8 grains
1 centigramme	= .01 ,,	= .15432348 8 ,,
1 grain	= .06479895 ,,	= 1. ,,
1 decigramme	= .1 ,,	= 1.54323488 ,,
1 gramme	= 1. ,,	= 15.4323488 ,,
1 décagramme	= 10. ,,	= 154.323488 ,,
1 ounce avoirdupois	= 28.34954 ,,	= 437.5 ,,
1 ounce troy	= 31.103496 ,,	= 480. ,,
1 hectogramme	= 100 ,,	= 1543.23488 ,,
1 pound troy	= 373.241955 ,,	= 5760. ,,

1 pound avoirdupois	=	453.59265 grms.	=	7000. grains.
1 kilogramme	=	1000 ,,	=	2.20462126 lbs.
1 myriagramme	=	10000 ,,	=	22.0462126.
1 cwt.	=	50.8023769 ,, kilos	=	112. lbs.
1 ton	=	1016.0475377 ,, kilos	=	2240. lbs.

NOTE.—The decimals are given extended, according to calculations, but they may be taken as short as is considered necessary for the purpose required, the last decimal being raised 1 if succeeded by a 5 or higher number.

The decimals have not always been repeated in the English equivalents, as it only requires care in placing the decimal point according to their grade, thus—

1 mètre - - - -	=	3.2808992 feet.
1 myriamètre, or 10,000 mètres	=	32808.992 feet.

CHAPTER XIII.

Questions on Part II.

[Some of the answers to these questions are worked out in full in order to show the best method of working calculations by the use of formulæ; and those questions so answered are marked with an asterisk (*). They are arranged in order after the list of questions.]

1. What is an atom defined to be? and show from the atomic hypothesis that $H\frac{1}{2}$ is incorrect.

2. Why, in the formula for water, must 2 atoms of hydrogen be represented?

3. From the consideration of the composition of water, what property do you attribute to the atom of oxygen besides its weight? Mention one or two other compounds of oxygen whose composition confirms this property.

4. Write down the formulæ of one or two bodies which can be formed by replacing the hydrogen in water by other elements.

5. What is a molecule defined to be? Write down the formulæ for a molecule of hydric nitrate, hydric sulphate, hydrogen, chlorine, phosphorus, nitric oxide, and sulphuric acid.

6. How is the weight of a molecule arrived at?

Write down the molecular weights of hydric chloride, hydric iodide, hydropotassic sulphate, and trihydric phosphate.

7. Write down two formulæ for the evolution of hydrogen from water.

8. Give an equation for the solution of hydrogen from hydric sulphate; compare fully the interchanges with those which take place when hydrogen is obtained from hydric chloride.

9. What is the difference between the products formed when hydric sulphate reacts with a metal, and with an oxide of a metal?

10. What is the action of red-hot iron on steam? Give the formula of the reaction.

11. Give in symbols the action of hydric sulphate on potassic dichromate. How much oxygen can be obtained from 105 grammes of potassic dichromate by the action of hydric sulphate? (*)

12. Given a kilogramme of manganic binoxide, how much oxygen can be obtained from it at a high temperature? (*)

13. How much carbon will 50 grammes of carbonic acid take up to form carbonic oxide? What volume of carbonic oxide will be formed? (*)

14. Give the formula for the formation of marsh gas from sodic acetate and sodic hydrate.

15. How many volumes of oxygen are required for the complete combustion of 150 litres of marsh gas? What will be the volume of carbonic acid formed? and what the weight of water? (*)

16. Give the formula for Dutch liquid, and state what volume its molecule occupies.

17. A litre of olefiant gas burns in two minutes; how much oxygen will be consumed in an hour if the burning continue at the same rate? what volume of carbonic acid will be found? and how much lime w uld be required to absorb it? (*)

18. Illustrate the law of multiple proportions, and give examples of it from the compounds of oxygen with chlorine.

19. Write down the equation for the formation of nitric oxide from hydric nitrate, silver being used to act upon the hydric nitrate.

20. Write down the equations for the formation of the following bodies:—

Ammonic nitrate from ammonia and hydric nitrate.

Ammonic sulphate from ammonia and hydric sulphate.

Cupric nitrate from copper and hydric nitrate.

Potassic sulphate from potash and hydric sulphate.

Potassic nitrate from potash and hydric nitrate.

21. What objection is there to writing the molecule of nitric oxide N_2O_2?

22. How much hydric nitrate must be decomposed to form nitric oxide sufficient to oxidise 19.2 grammes of sulphurous acid into sulphuric? (*)

23. Give the various proportions in which oxygen combines with nitrogen. Name the compounds in each case, and give the symbols for their molecules. What laws do you observe governing these combinations?

24. Which body contains most oxygen, phosphoric or phosphorous acid? What is the name for the hydrogen salt of phosphoric acid, and what of the hydrogen salt of phosphorous acid?

25. Describe the preparation of nitrous oxide, and give equation.

26. Can nitrous oxide be liquefied? At what pressure?

27. What is the density of nitrous oxide?

28. What are the physical properties of nitrous oxide? What effect has it on the human system?

29. Does nitrous oxide support combustion?

30. What volume of nitrous oxide and what weight of water are produced by the decomposition of 560 grammes of ammonic nitrate?—*Ans.* 156.8 litres and 252 grammes of water.

31. What weight and volume of nitrous oxide can be obtained from 880 grammes of ammonic nitrate?—*Ans.* 484 grammes and 246.4 litres.

32. It is desired to make 4.4 grammes of nitrous oxide, what weight of ammonic nitrate must be taken?—*Ans.* 8 grammes.

33. What weight of ammonic nitrate must be taken to produce 10.8 grammes of water and 6.72 litres of nitrous oxide?—*Ans.* 24 grammes.

34. What volumes of oxygen and nitrogen in the free state go to form 672 litres of nitrous oxide?—*Ans.* 672 of nitrogen and 336 of oxygen.

35. What volume of hydrogen can be burnt by 132 grammes of nitrous oxide?—*Ans.* 67.2 litres.

36. What volume of nitrous oxide is requisite for the combustion of 20 grammes of hydrogen?—*Ans.* 224 litres.

37. Under what circumstances does nitric oxide support combustion?

38. Can nitric oxide be condensed to a liquid? If so, at what temperature?

39. How is nitric oxide usually prepared? Express the reaction by formulæ.

40. Give the composition and density of nitric oxide.

41. What is the volume of 360 grammes of nitric oxide?—*Ans.* 268.8 litres.

42. What is the weight of 224 litres of nitric oxide? —*Ans.* 300 grammes.

43. 180 grammes of nitric oxide are required, what weights of hydric nitrate and copper must be used, and what will be the volume of the nitric oxide evolved? —*Ans.* 1.512 kilogrammes of hydric nitrate; 571.5 grammes of copper; 134.4 litres of nitric oxide.

44. When 252 grammes of hydric nitrate act on metallic copper, what weights of the following substances are found, viz. nitric oxide, water, and cupric nitrate? —*Ans.* 30 grammes of nitric oxide, 36 grammes of water, and 281.25 grammes of cupric nitrate.

45. What weight of copper and hydric nitrate must be taken to produce 22.4 litres of nitric oxide?—*Ans.* 95.25 grammes of copper and 252 grammes of hydric nitrate.

46. When oxygen unites with nitric oxide what is the product? What are the physical properties of the gas? What is its colour in the liquid state?

47. When nitrous acid stands in a vessel over water, what changes take place? Represent the change by an equation.

48. What is the composition of nitrous acid? What its density?

49. What volume of oxygen must be added to 180 grammes of nitric oxide to convert it into nitrous acid, and what will be the volume of the nitrous acid formed?

—*Ans.* 33.6 litres of oxygen and 67.2 litres of nitrous acid will be formed.

50. At what temperature does nitrous acid assume the gaseous from the liquid state?

51. What are the products of the decomposition of plumbic nitrate? Give the equation.

52. What is the colour of liquid nitric peroxide? At what temperature is it condensed to a liquid? What are its physical properties in the gaseous state?

53. What is the density of nitric peroxide? What is its composition?

54. What are the weights of the two gases formed by the decomposition of 993 grammes of plumbic nitrate by heat? Which of the two can be condensed to a liquid?—*Ans.* Oxygen 48 grammes, nitric peroxide 276 grammes.

55. What weight of plumbic nitrate must be taken to produce 46 grammes of nitric peroxide?—*Ans.* 165.5 grammes.

56. What are the weights of the three products resulting from the decomposition of 496.5 grammes of plumbic nitrate?—*Ans.* 334.5 grammes of plumbic oxide, 24 grammes of oxygen, and 138 grammes of nitric peroxide.

57. By what other names has nitric acid been called?

58. In what form is nitric acid obtained when chlorine is passed over argentic nitrate? Give the equation for the reaction.

59. How would you determine the relative proportions of oxygen and nitrogen in nitric acid?

60. What is meant by the water type? Write out on this type zinc sulphate, and show its formation from

zinc and hydric sulphate, also sodic nitrate and calcic nitrate.

61. Write down the equation for obtaining chlorine from magnesic chloride by the action on it of hydric sulphate and manganic binoxide? How much chlorine can be obtained from magnesic chloride by the action of 15 grammes of manganic binoxide? (*)

62. What volume do 177.5 grammes of chlorine occupy? and what weight of hydrogen must be added to produce a mixture in the proportions for making hydric chloride?—*Ans.* 56 litres of chlorine and 5 grammes of hydrogen. (*)

63. 20 grammes of iodine are suspended in water, and 1 litre of hydric sulphide is passed through the mixture; find the weights of all the products. (*)

Answers to some of the preceding Questions.

No. 11.

$$2(K_2Cr_2O_7) + 8(H_2SO_4) = 2(K_2SO_4) + 2\{Cr_2(SO_4)_3\}$$
2(295)

$$+ 8(H_2O) + 3O_2$$
3(32)

$$K_2 = 78$$
$$Cr_2 = 105$$
$$O_7 = 112$$
$$\overline{295}$$

2(295) grms. of potassic dichromate give 3(32) grms of oxygen

$\therefore$ 1 gramme of potassic dichromate gives $\frac{48}{295}$ ditto

$\therefore$ 105 grms. ditto give $\frac{48\times105}{295} =$ 17.0847 grms. of oxygen.

No. 12.

$$3(MnO_2) = Mn_3O_4 + O_2$$
3(87) 32

$$Mn = 55$$
$$O_2 = 32$$
$$\overline{87}$$

261 grms. of manganic binoxide give 32 grms. of oxygen

$\therefore$ 1 grm. gives $\frac{32}{261}$ grms. of oxygen

$\therefore$ 1000 give $\frac{32\times1000}{261} = 122.605$ grms. of oxygen.

No. 13.

$$\begin{array}{llcccc} C = 12 & & CO_2 & + \ C & = & 2(CO) \\ O_2 = 32 & & 44 & \ \ 12 & & 2(28) \\ \overline{\ \ \ 44} & & \text{2 vols. (22.4 litres)} & & & \text{4 vols. (44.8 litres)} \end{array}$$

44 grms. take up 12 grms. of carbon to form 56 grms. of carbonic oxide,

$$\therefore \text{ 1 grm. requires } \frac{12}{44}$$

$$\therefore \text{ 50 grms. require } \frac{12 \times 50}{44} = 13.6\dot{3}6 \text{ grms. of carbon.}$$

56 grms. of carbonic oxide occupy 44.8 litres

∴ 14 grms. ,, ,, 11.2 litres

$$\therefore \text{ 1 grm. occupies } \frac{11.2}{14} \text{ litres}$$

$$\therefore \text{ 13.636 grms. occupy } \frac{11.2 \times 13.6\dot{3}6}{14} = 10.9088 \text{ litres.}$$

No. 15.

$$\begin{array}{llccccccc} C = 12 & & CH_4 & + & O_4 & = & CO_2 & + & 2(H_2O) \\ H_4 = \ \ 4 & & 16 & & & & & & 2(18) \\ \overline{\ \ \ 16} & & \text{2 vols.} & & \text{4 vols.} & & \text{2 vols.} & & \text{4 vols.} \end{array}$$

∴ as 2 vols. of marsh gas require 4 vols. of oxygen for their combustion, 150 litres will require 300 litres of oxygen; and as the volume of carbonic acid is the same as that of the marsh gas, it will be 150 litres.

2(11.2) litres of marsh gas on combustion yield 36 grammes of water,

$$\therefore \text{ 1 litre yields } \frac{36}{22.4}$$

$$\therefore \text{ 150 litres yield } \frac{36 \times 150}{22.4} = 241.071 \text{ grms. of water.}$$

No. 17.

1 litre of olefiant gas burns in 2 minutes,
∴ 30 litres burn in an hour.

$$\begin{array}{ccccccccc} C_2H_4 & + & 3O_2 & = & 2(CO_2) & + & 2(H_2O) \\ \text{2 vols.} & & \text{6 vols.} & & \text{4 vols.} & & \text{4 vols.} \end{array}$$

∴ 30 litres require 90 litres of oxygen for their combustion, and 60 litres of carbonic acid are formed.

$$\begin{array}{lllllll} Ca = 40 & \quad & CaO & + & CO_2 & = & CaCO_3 \\ O\ = 16 & & 56 & & \text{22.4 litres} & & \\ \overline{\quad 56} & & & & & & \end{array}$$

22.4 litres of carbonic acid require 56 grms. of lime

$$\text{1 litre requires } \frac{56}{22.4}$$

$$\therefore \text{ 60 litres require } \frac{56 \times 60}{22.4} = \text{150 grms. of lime.}$$

No. 22.

$$SO_2 + O = SO_3$$

$$3Cu + 8(HNO_3) = 3\{Cu(NO_3)_2\} + 4(H_2O) + 2(NO)$$

2(NO) takes up 1 atom of oxygen, which it gives up to the molecule of sulphurous acid, SO_2, to form a molecule of sulphuric acid, SO_3; and 2(NO) requires $8(HNO_3)$ for its production $8(HNO_3) = 504$.

$$\begin{array}{l} S\ = 32 \\ O_2 = 32 \\ \overline{\quad 64} \end{array}$$

64 grms. of sulphurous acid require 504 grms. of hydric nitrate

$$\therefore \text{ 1 grm. requires } \frac{504}{64}$$

$$\therefore \text{ 19.2 grms. require } \frac{504 \times 19.2}{64} = \text{151.2 grms. of hydric nitrate.}$$

No. 61.

$$\underset{87}{MnO_2} + MgCl_2 + 2(H_2SO_4) = MgSO_4 + MnSO_4 + 2(H_2O) + \underset{71}{Cl_2}$$

$Mn = 55$
$O_2 = 32$
87

87 grms. of manganic binoxide set free 71 grms. of chlorine

$\therefore$ 1 grm. sets free $\frac{71}{87}$

$$\therefore 15 \text{ grms. set free } \frac{71 \times 15}{87} = 12.241 \text{ grms. of chlorine.}$$

No. 62.

35.5 grms. of chlorine occupy 11.2 litres

$\therefore$ 1 grm. occupies $\frac{11.2}{35.5}$

$$\therefore 177.5 \text{ grms. occupy } \frac{11.2 \times 177.5}{35.5} = 56 \text{ litres}$$

$HCl = 2$ vols., 1 vol. of hydrogen and 1 vol. of chlorine, $\therefore$ 56 litres of chlorine require 56 litres of hydrogen to form 112 litres of hydric chloride.

11.2 litres of hydrogen weigh 1 gramme

$\therefore$ 56 litres weigh 5 grammes,

the weight of hydrogen required is 5 grammes.

No. 63.

1 litre of hydric sulphide weighs $\frac{34}{22.4}$ grms. $= 1.517$ grms.

$$\begin{array}{ccccccccc} H_2S & + & I_2 & = & S & + & 2HI \\ 34 & & 254 & & 32 & & 256 \end{array}$$

254 grms. of iodine require 34 grms. of hydric sulphide

$$\therefore 1 \text{ grm. requires } \frac{254}{34}$$

and 1.517 grms. require $\frac{254 \times 1.517}{34} = 11.33\dot{3}$ grms. of iodine required to unite with the hydrogen of 1 litre of hydric sulphide.

But 20 grms. of iodine are in suspension,

$\therefore 20 - 11.332 = 8.668$ of iodine remain uncombined.

34 grms. of hydric sulphide would deposit 32 grms. of sulphur,

$$\therefore 1 \text{ grm. would deposit } \frac{32}{34}$$

and 1.517 grms. would deposit $\frac{32 \times 1.517}{34} =$

1.4277 grms. of sulphur deposited.

34 grms. of hydric sulphide can form with iodine 256 grms. of hydric iodide,

$\therefore$ 1 grm. of hydric sulphide can form $\frac{256}{34}$

and 1.517 grms. ,, ,, $\frac{256 \times 1.517}{34} =$

11.4221 grms. of hydric iodide.

CHAPTER XIV.

ADVICE TO CANDIDATES UNDER EXAMINATION.

The object of an examination is to discover what the candidate knows about his subject, and as the examination is limited to a short time, it is only possible for the examiner to put a few questions; these he selects to the best of his judgment from different parts of the subject in order to test the knowledge of the candidate, who is expected to answer them precisely, and not to bring in irrelevant matter. The examiner has only about eight chances of finding out a candidate's knowledge, for only about eight questions are asked, and these he expects to be answered. No knowledge, however correct, if it does not bear directly on the questions, can be taken into account. Moreover, when a person writes very little about the question, and very much which is foreign to it, the conclusion is that he knows but little about his subject, and is trying to hide his ignorance by a show of knowledge, and this is *most prejudicial* to his interests. For example, a question is asked, How is oxygen prepared from manganic binoxide? The candidate does not know, but he writes fully about its preparation from potassic chlorate, describes its properties, and dilates on the brilliancy of the combustion in it of

phosphorus, &c. Such an answer is not what the examiner wants; he sees that the candidate is ignorant of what he expects him to know, and the result is that he gives no marks for his answer. Again, some persons seem to think that it is necessary to write *long* answers; of course the answers must be sufficiently long to express all that the question requires, but the fewer words used beyond this the better; for looking over examination papers is not so pleasant a task that an examiner would wish it to be unnecessarily prolonged. Moreover, *clearness* is of great importance. Doubtless many an answer which has contained much correct matter has been marked low, or received no marks at all, because the examiner could not spare the time to unravel the mystery in which the candidate had thought well to involve it. To return to the question just asked,—How is oxygen prepared from manganic binoxide?—*Ans.* Manganic binoxide is placed in an iron tube closed at one end, the other being connected with a delivery tube, or in a fireclay retort to which a tube is fixed, to conduct away the gas as it is evolved. Whichever of these vessels may be employed, it is heated to a high temperature, the oxygen passes through the delivery tube, and is best collected over water.

There is no need to say more than this—the question does not require it. Had the question been, Explain *fully* the manufacture of oxygen from manganic binoxide, then it is clear that you would be expected to enter into other matters,—such as the quantity of oxygen evolved, &c., and the answer must be worded in some such way as this: Manganic binoxide only parts with one-third of its oxygen at a white heat; therefore,

if a given quantity of oxygen be required, say 16 grammes, 130.5 grammes of binoxide must be taken, according to the equation $3(MnO_2) = Mn_3O_4 + O_2$, the binoxide must be heated in an iron or fireclay vessel and the gas collected over water.

When you first look at the examination paper you will perhaps feel that there is not a question in it which you can answer. This is the result of nervousness. Read it over carefully once or twice, and gradually you will find the light dawn upon you. When you have found that you can do something, take the question you know most about, and answer it as well and as carefully as you can, and proceed in this way till you find you can do no more. Your success in your first answer will give you confidence. You may next apply yourself seriously to the more difficult questions, and then, with your mind sharpened by its previous exercise, you will be able to understand and explain what at first appeared to be altogether beyond you. There is no more constant source of failure in examinations than the attempt to answer the most difficult questions first—you think and puzzle your brains till they get absolutely muddled, and when you give the question up in despair, and turn to others more easy, which you could have answered well had you undertaken them first, you find that you are uncertain about what to say; you do not remember whether such an action takes place in this case or whether it belongs to some other bodies, and eventually you write down something as well as you are able, and from the time you leave the examination room till the publication of the list, you are calling to mind your answers—what you said and what you might have

said, had you adopted a different course. The remarks here made apply only to those who have conscientiously studied their subject. The success of others who have "crammed," is always a matter of chance—rules, therefore, can have no reference to them.

If a person has learnt his Chemistry only from books, or if he has got it up by learning answers to questions, it is clear that having no knowledge of the principles on which reactions take place, he cannot answer a question which requires a knowledge of those principles. Such a question as, How is hydric chloride made? might be answered from simple book-knowledge; but if the same question were put in this way: Sodic chloride is heated with an acid liquid, which can exchange its hydrogen for sodium, under the influence of heat, what gas will be given off?—it could not be answered by any but those who have studied their subject, and made themselves acquainted with the first principles of it. Now, inasmuch as the Chemistry papers of the London University Examiners show a desire to suppress the cram system, it behoves all candidates to abandon it, if they wish to ensure success at the examination, if they are actuated by no higher motive.

Frequently the gist of a question is overlooked, and the answer given, although bearing on the subject, does not fully answer it. This question will be found in the London University Matriculation papers given at the end of this chapter: "Explain how you would prove the presence of water, carbonic acid, oxygen, and nitrogen in atmospheric air?" This question is sometimes worded, How do you *obtain* oxygen, &c.,

from air? At first sight it appears that a description of the ordinary method of analysis of air, if given, would be sufficient. The question, however, requires *proof* of the existence of certain substances in air, and we can only arrive at this proof by examining each substance by its special characteristics. We can see water; therefore, if we collect it from air, so that we can see it, and apply other tests to it, we may assert that it is present in it. Now, to collect water from air, we have only to bring into it a substance, such as a tumbler of ice, which is colder than the air, and water, as such, will be condensed upon its surface. Carbonic acid we cannot see, but many of its compounds are very visible; one, calcic carbonate, is well known. If then lime water be exposed to air, white calcic carbonate is formed and precipitated. Or if air be passed into a solution of sodic hydrate, sodic carbonate is formed; but here we cannot see any difference in the liquid before and after the passage of the air, but, on addition of an acid liquid, effervescence takes place and carbonic acid is set free; this we can collect in a vessel, and test by a burning splint, which it will extinguish; and with lime water, the lime of which it will convert into insoluble calcic carbonate. The presence of oxygen can be proved by heating mercury or barytic oxide in air; oxygen will be taken up by these substances, and may be set free, collected and tested by its re-kindling a glowing splint, or by absorption by potassic pyrogallate. The presence of nitrogen in air can be proved by burning phosphorus in a vessel of pure air; the oxygen will be taken by the phosphorus, and nitrogen will be left. which can be

tested by its extinguishing a burning candle, and by its not forming any white precipitate with lime water, and by its not burning.

Again, in the following question, "A current of pure air passes through a bright coke fire, what gases are contained in the air beyond the fire?"—you are not required by this question to describe what takes place *in* the lower part or middle of the fire, as you would be if the question was, Explain the burning of a common fire; you are simply asked what gases are contained in the air *beyond* the fire. Now these are the gases usually contained in air, and the gases formed by the combustion of the fire at its *upper* part—nitrogen and oxygen in the air, carbonic acid, steam or water gas in the immediate neighbourhood of the fire, *i.e.* before its condensation. A question was asked in the June paper, 1868—"During the Abyssinian expedition, drinkable water had to be prepared from sea water; show how this can be done." Here it would not be enough to state that distillation renders the water fresh by ridding it of its saline constituents; sea water, so purified, is not *drinkable.* It is necessary to state that after distillation it should be aërated by filtration through animal charcoal. Sufficient illustrations have been given to show the importance of carefully thinking over the exact scope of the question, and of discovering what it is the examiners require. The few hints given in this chapter will, it is hoped, prove useful.

After the list of University Examination Questions, answers to those which involve calculations are given, and the method of working them out is explained.

CHAPTER XV.

LONDON UNIVERSITY EXAMINATION PAPERS.

Matriculation.

Questions on Chemistry from Jan. 1865 to Jan. 1869, inclusive.

1. Describe some of the most important natural substances in which oxygen is contained: and explain the preparation of the gas from the black oxide of manganese. Name the substances which are formed when carbon, hydrogen, phosphorus, and sulphur respectively are burned in an excess of oxygen.

2. Four bottles containing oxygen, hydrogen, nitrogen, and chlorine respectively are given to you, with the request to determine the nature of the gas in each bottle. How would you distinguish each of these gases from the other?

3. Illustrate the law of multiple proportions by a series of examples including compounds of carbon with oxygen.

4. How would you prove that carbonic acid consists of carbon and oxygen? How is carbonic acid usually prepared? How do you distinguish carbonic acid from nitrogen?

5. Describe the chief processes by which oxygen is

being removed from the atmosphere. How is it that the ratio of oxygen to nitrogen is nevertheless practically invariable?

6. How much oxygen is contained in 10,000 grammes of potassic dichromate $K_2Cr_2O_7$, $K = 39$, $O = 16$, $Cr. = 52.5$? How much oxygen is given off when the same quantity of the salt is treated with an excess of oil of vitriol?

7. In what proportion by volume does oxygen combine with each of the following gases—viz. hydrogen, carbonic oxide, marsh gas, and olefiant gas?

8. What degree of Centigrade corresponds to $-40°$F.? (Give the details of the calculation.) What degree Fahr. corresponds to 35.°C.?

9. Why is atmospheric air supposed to be a mechanical mixture, and not a chemical compound of oxygen and nitrogen?

10. Give the molecular weights of common salt, iodide of potassium, carbonic acid, nitrate of silver (following atomic weights being adopted—chlorine 35.5, sodium 23, iodine 127, potassium 39, carbon 12, silver 108, oxygen 16, nitrogen 14).

11. Write in symbols the following compounds—water, sulphuric acid, nitric acid, nitrous oxide, sulphuretted hydrogen.

12. Explain the following chemical reactions—

(a) The decomposition of steam by red-hot iron.

(b) The decomposition of water by potassium in presence and absence of atmospheric air.

(c) The formation of hydrochloric acid and sulphate of soda from common salt and sulphuric acid.

(d) The action of chlorine on sulphuretted hydrogen.

13. What is ammonia? How is it obtained? Give its composition by weight and by volume.

14. Explain why a candle cannot burn long in a closed bottle. How is the composition of dry air affected by its passage over red-hot iron or copper? How much air is needed for the complete combustion of 6 grammes of pure carbon?

15. Describe an experiment by which you could prove that silica must be classed among acids.

16. Describe the preparation of pure hydrogen. How would you prove that hydrogen is lighter than air?

17. How many grammes of oxygen and how many of hydrogen are required in order to prepare 500 grammes of water? How many litres will the hydrogen measure if each gramme measures 11.2 litres? How many litres will the oxygen measure, its density being 16 times as great as that of hydrogen?

18. Explain how you would prove the presence of water, carbonic acid, oxygen and nitrogen in atmospheric air?

19. Explain fully why the flame of a common gas-light is brighter than the flame of a Bunsen burner. (A mixture of air and gas is burnt in the latter case.)

20. How is pure carbonic acid prepared? How much heavier is it than hydrogen?

21. Explain fully the meaning of the following symbols:—O_2, H_2, Cl_2, Br_2, I_2, C, N_2, S, P, CO_2, SO_3, H_2O, and NH_3.

22. Describe experiments by which you would endeavour to ascertain whether a given substance (dissolved in water) is an acid, a base, or a salt. Give an example of a body belonging to each of the above-

named classes, and enumerate those properties of each body which you consider to afford the best proof that it belongs to the class to which you assign it.

23. How many volumes of their respective elements are contained in 2 volumes of each of the following gases—hydrochloric acid, steam, and ammonia?

24. Describe an apparatus for the decomposition of water by the galvanic current, and the result of the decomposition.

25. Name the constituents of the atmosphere. How did Lavoisier show the presence of oxygen in air?

26. How is red phosphorus made from clear phosphorus? Describe the chief differences between the two varieties. Why is common phosphorus kept under water?

27. Describe the products which are found during the combustion of coal gas, and explain what means you would adopt for their detection. Why does olefiant gas burn with a brighter flame than hydrogen?

28. How would you prepare the following gases—hydrogen, chlorine, nitrogen? Give in each case an explanation of your process.

29. Describe how you would distinguish oxygen from common air. What class of bodies do we call oxides? Name some oxides of an acid nature, and some basic oxides.

30. Describe how you would prepare hydrogen from water without the assistance of an acid, and explain your process. How can a mixture of 2 volumes of hydrogen and 1 of oxygen be converted into water?

31. In what respects does rain-water differ from sea water? Enumerate the most common constituents of

spring water. How is pure water obtained from spring water? and how would you ascertain whether a given sample of spring water contains sulphuretted hydrogen?

32. Describe the structure of the flame of a common candle, and explain why the flame loses in brightness when a current of common air is blown into it.

33. Explain why oxygen, hydrogen, nitrogen, and carbon are considered elements. Describe one experiment by which you could prove that diamond and graphite consist of the same element.

34. Describe experiments by which you could distinguish a mechanical mixture of oxygen and hydrogen from the compound of those elements.

35. 100 volumes of nitrogen are given. How many volumes of oxygen must be added in order to obtain a mixture of the same composition as pure atmospheric air? What volume of oxygen is needed for the complete combustion of 1 gramme of carbon? (16 grammes of oxygen measure 11.2 litres.)

36. Describe and explain the preparation of oxygen from potassic chlorate (chlorate of potash); and name the substances which are formed when carbon, hydrogen, sulphur, and phosphorus are respectively burned in excess of oxygen. Which of the following bodies can be burnt in oxygen—sulphur, nitrogen, coke, sulphurous acid, wax, iodine, chlorine, ammonia?

37. What weight of oil of vitriol ($H_2 SO_4$) must be decomposed for the evolution of 1 gramme of hydrogen?

38. Give the symbols of the following elements and compounds—hydrogen, nitrogen, nitrous oxide, carbon, marsh gas, chlorine, bromine, phosphorus, phosphoric acid, and silicic acid.

39. Describe the effects produced by chlorine on the following substances—indigo solution, copper foil, powdered antimony, powdered charcoal, phosphorus, and sulphur.

40. How much carbon could be completely burnt by a ton of air?

41. Explain by means of symbols the preparation of hydrochloric acid from common salt and sulphuric acid. How many volumes of chlorine and hydrogen are required for the preparation of 2 volumes of hydrochloric acid? How does water act on hydrochloric acid gas?

42. What degree of the Centigrade scale corresponds to 85°F.? and what degree of the Fahrenheit scale to —40°C.?

43. Name the most important natural substance in which nitrogen is contained; and describe how you would prepare nitrogen. Enumerate the principal properties of nitrogen; and show how it may be distinguished from oxygen, hydrogen, and carbonic acid.

44. Name the two chief constituents of our atmosphere: state in what proportion they occur in it: and give their respective symbols. Why is air considered to be a mechanical mixture, and not a chemical compound of these constituents?

45. Describe how you would prepare chlorine, and explain your process. Also give an outline of the chief chemical and physical properties of this element.

46. Two liquids are given, one containing an acid, and the other a base. Describe the experiments you would perform in order to ascertain which of the two liquids contains the acid.

47. How much anhydrous phosphoric acid is formed by the combustion of 1 gramme of phosphorus ($P = 31$)? How much pure air is needed for the production of this phosphoric acid?

48. What weight of carbon is needed for the formation of 88 grammes of carbonic acid? What is the volume of this weight of acid at 0°Centigrade and 30 inches barometric pressure?

49. Describe how you would prepare carbonic acid from chalk; and explain your process by means of symbols. How would you test whether the air in a well consists of ordinary air or of carbonic acid? By what means would you detect carbonic acid in spring water?

50. It is asserted that 1 molecule of hydrochloric acid consists of 1 atom of chlorine and 1 atom of hydrogen; and 1 molecule of ammonia of 1 atom of nitrogen and 3 atoms of hydrogen. What do we mean by an atom of nitrogen and an atom of hydrogen? Calculate the per-centage composition of ammonia.

51. During the Abyssinian Expedition drinkable water had to be prepared from sea-water. Show how this can be done. Enumerate the principal constituents of sea-water.

52. A current of pure air passes through a bright coke fire. What gases are contained in the air beyond the fire? Give a short description of each of them.

53. Describe the chief allotropic modifications of sulphur; and name the chief substances from which sulphur is obtained. How is sulphurous acid obtained? How much sulphur and how much oxygen are contained in 100 grammes of sulphurous acid?

54. Describe the preparation of ammonia from ammonic chloride (sal ammoniac); and explain your process by means of symbols. What volume of hydrochloric acid gas is needed in order to condense a cubic inch of ammonia?

55. How much pure air ought to be mixed with 100 measures of fire-damp (marsh gas) in a coal-mine, in order to supply materials for the complete combustion of the marsh gas during explosion?

56. Describe how you would prepare hydrogen from water, zinc, and sulphuric acid; and explain your process by means of symbols. Sketch the leading properties of hydrogen; and enumerate native compounds of this element.

57. Give the formulæ of the following substances: water, ammonia, silica, carbonic acid, sulphuric acid, marsh gas, and phosphuretted hydrogen.

58. Enumerate the chief native sources of carbonic acid; and mention the causes which tend to keep down the amount of this gas in our atmosphere.

59. Describe the preparation of aqueous hydrochloric acid; and explain the process by an equation. How would you distinguish it from an aqueous solution of nitric acid?

60. A sample of vinegar is supposed to be adulterated with sulphuric acid. How would you ascertain the presence of sulphuric acid in vinegar?

61. Describe how you would render silica soluble in hydrochloric acid, and how the soluble modification of this substance can be converted into the insoluble form.

Answers to Calculations in Matriculation Examination Papers.

No. 6.

The formula for potassic dichromate is $K_2Cr_2O_7$ $K_2 = 2(39) = 78$; $Cr_2 = 2(52.5) = 105$; $O_7 = 7(16) = 112$; therefore the molecular weight of $K_2Cr_2O_7 = 295$. 295 grammes of potassic dichromate contain 112 grammes of oxygen.

	Grms. of $K_2Cr_2O_7$.		Grms. of $K_2Cr_2O_7$.		Grms. of oxygen.		Grms. of oxygen.
$\therefore$	295	:	10,000	: :	112	:	x

$$x = \frac{10.000 \times 112}{295} = 3796.61 \text{ grms. of oxygen.}$$

When potassic dichromate is heated with excess of hydric sulphate, the following equation represents the reaction—

$$\underset{295}{K_2Cr_2O_7} + 5(H_2SO_4) = Cr_2(SO_4)_3 + 2(KHSO_4) + 4(H_2O) + \underset{48}{O_3}$$

295 grammes of $K_2Cr_2O_7$ yield 48 grammes of oxygen,

$$\therefore \; 295 : 10,000 :: 48 : x$$

$$x = \frac{10,000 \times 48}{295} = 1627.118 \text{ grms. of oxygen.}$$

Therefore 10,000 grammes of $K_2Cr_2O_7$ contain 3796.61 grammes of oxygen, and when heated with H_2SO_4 in excess they give off 1627.118 grammes of oxygen.

No. 7.

1 volume of oxygen combines with 2 volumes of hydrogen to form water.

1 volume of oxygen combines with 2 volumes of carbonic oxide to form carbonic acid.

4 volumes of oxygen combine with 2 volumes of marsh gas to form water and carbonic acid.

6 volumes of oxygen combine with 2 volumes of olefiant gas to form water and carbonic acid.

No. 8.

32° Fahrenheit corresponds to zero Centigrade, and —40°F. are 40 degrees below zero F.; therefore—40°F. is 40° + 32°, *i.e.* 72°F. below freezing point. Now, Centigrade degrees are to Fahrenheit degrees as 5 to 9; therefore $\frac{72 \times 5}{9}$ = 40 degrees below zero Centigrade, or — 40°C.

Or the calculation may be worked by the formula

$$C = \frac{(F - 32)5}{9}$$

$$C = \frac{(-40 - 32)5}{9}$$

$$= \frac{-72 \times 5}{9} = -40$$

Fahrenheit degrees are to Centigrade degrees as 9 to 5, therefore $\frac{35 \times 9}{5} = 63$; and as 0°C. corresponds to 32°F. above zero, therefore we must add 32 to 63, which gives 95°F., or by the formula

$$F = \frac{C9}{5} + 32$$

$$= \frac{35 \times 9}{5} + 32$$

$$= 63 + 32 = 95°F.$$

No. 17.

16 grammes of oxygen combine with 2 grammes of hydrogen to form 18 grammes of water.

	Grms. of water.		Grms. of water.		Grms. of oxygen.		Grms. of oxygen.
$\therefore$	18	:	500	: :	16	:	x

$$x = \frac{500 \times 16}{18} = 444.\dot{4} \text{ grms. of oxygen.}$$

Grms. of water.		Grms. of oxygen.	
500	—	$444.\dot{4}$	$= 55.\dot{5}$ grms. of hydrogen.

The following proportions will give the volume of the oxygen :—

Grms. of oxygen.		Grms. of oxygen.		Litres.		Litres.
16	:	$444.\dot{4}$	: :	11.2	:	x

$$x = \frac{444.\dot{4} \times 11.2}{16} = 311.08 \text{ litres of oxygen.}$$

And if we multiply 311.08 by 2 it will of course give the volume of the hydrogen—

$$311.08 \times 2 = 622.16 \text{ litres of hydrogen.}$$

No. 23.

Hydrochloric acid $HCl = 2$ volumes, 1 volume of chlorine and 1 volume of hydrogen.

Steam $H_2O = 2$ volumes, 2 volumes of hydrogen and 1 volume of oxygen.

Ammonia $NH_3 = 2$ volumes, 1 volume of nitrogen and 3 volumes of hydrogen.

No. 35.

In 100 volumes of air there are 21 volumes of oxygen and 79 volumes of nitrogen.

$\therefore\ 79 : 100 :: 21 : x$

$$x = \frac{100 \times 21}{79} = 26.582 \text{ vols. of oxygen.}$$

100 volumes of nitrogen require 26.582 volumes of oxygen.

12 grammes of carbon unite with 32 grammes of oxygen to form 44 grammes of (CO_2) carbonic acid. O_2 = 32 grammes = 22.4 litres.

$12 : 1 :: 22.4 : x$

$$x = \frac{22.4}{12} = 1.8\dot{6} \text{ litres of oxygen.}$$

1 gramme of carbon requires for its complete combustion $1.8\dot{6}$ litre of oxygen.

No. 37.

The following equation represents the decomposition—

$$\underset{98}{H_2SO_4} + Zn = Zn.SO_4 + \underset{2}{H_2}$$

98 parts by weight of oil of vitriol give off, when decomposed, 2 parts by weight of hydrogen.

$\therefore$ 49 grammes will require to be decomposed for the evolution of 1 gramme of hydrogen.

No. 40.

Every 100 tons of air contain 23.1 tons of oxygen; $\therefore$ 1 ton of air contains .231 ton of oxygen; and 32 tons of oxygen are sufficient for the complete combustion of 12 tons of carbon, and the weight of carbon that can be burnt by .231 ton of oxygen is found as follows:—

Tons of oxygen.		Ton of oxygen.		Tons of carbon.		Tons of carbon.
32	:	.231	: :	12	:	x

$$x = \frac{.231 \times 12}{32} = .086625$$

1 ton of air is sufficient for the complete combustion of .086625 of a ton of carbon.

No. 42.

$$C = \frac{(F-32)5}{9}$$

$$= \frac{(85-32)5}{9} = \frac{265}{9} = 29.\dot{4}$$

85°F. corresponds to 29.$\dot{4}$°C.

$$F = \frac{C \times 9}{5} + 32$$

$$= \frac{-40 \times 9}{5} + 32 = -40°F.$$

No. 47.

2 atoms of phosphorus, weighing together 62 grammes, require for their combustion 5 atoms of oxygen, weighing together 80 grammes.

Grms. of phosphorus.		Grm. of phosphorus.		Grms. of oxygen.		Grms. of oxygen.
$\therefore$ 62	:	1	::	80	:	x

$$x = \frac{80}{62} = 1.3.$$

1.3 gramme of oxygen is required to burn 1 gramme of phosphorus. 100 grains of air contain 23 grammes of oxygen.

$$\therefore\ 23 : 1.3 :: 100 : x$$

$$x = \frac{100 \times 1.3}{23} = 5.65.$$

The air required for the combustion of 1 gramme of phosphorus is 5.65 grammes.

The weight of anhydrous phosphoric acid formed is equal to the sum of the two elements used for its formation.

	Phosphorus.		Oxygen.		
that is,	1	+	1.3	=	2.3 grammes.

No. 48.

12 parts by weight of carbon require 32 parts by weight of oxygen to form 44 parts by weight of carbonic acid.

	Grms. of CO_2		Grms. of CO_2		Grms. of carbon.		Grms. of carbon.
∴	44	:	88	::	12	:	x

$$x = \frac{88 \times 12}{44} = 24 \text{ grammes of carbon required.}$$

44 grammes of CO_2 occupy at 0°C., and 30 inches barometric pressure, 22.4 litres; ∴ 88 grammes of CO_2 occupy 2(22.4) litres, or 44.8 litres.

No. 50.

The molecule of ammonia $NH_3 = 17$ $\begin{array}{l} N = 14 \\ H_3 = 3 \end{array}$

So that 17 grammes of ammonia contain 14 grammes of nitrogen and 3 grammes of hydrogen. The percentage of hydrogen is found as follows:—

$$17 : 100 :: 3 : x$$

$$x = \frac{300}{17} = 17.647 \text{ grammes of hydrogen in 100 grms. of ammonia.}$$

∴ $100 - 17.647 = 82.353$ grammes of nitrogen in 100 grammes of ammonia.

82.353 parts of nitrogen
17.647 ditto of hydrogen

100.000

No. 54.

$$NH_3 + HCl = NH_4Cl.$$

Here 1 molecule of ammonia unites with 1 molecule of hydrochloric acid, and these 2 molecules occupy the same volume; $NH_3 = 2$ vols.; $HCl = 2$ vols.; and therefore 1 cubic inch of ammonia will require 1 cubic inch of hydrochloric acid to form *solid* chloride of ammonium, NH_4Cl.

No. 55.

The following is the equation which represents the combustion of a molecule of marsh gas or fire-damp :—

$$CH_4 + O_4 = CO_2 + 2(H_2O) \qquad CH_4 = 2 \text{ vols.}$$
$$O_4 = 4 \text{ vols.}$$

Marsh gas therefore requires double its volume of oxygen for its complete combustion; so that 100 measures, or volumes, will require 200 measures of oxygen. As air contains, in every 100 measures, 21 measures of oxygen, the following proportion will give the volume of air which contains 200 measures of oxygen :—

Vols. of oxygen.		Vols. of oxygen.		Vols. of air.		Vols. of air.
21	:	200	:	100	: :	x

$$x = \frac{200 \times 100}{21} = 952.381 \text{ vols. of air.}$$

952.381 measures of air are necessary for the complete combustion of 100 measures of fire-damp.

Matriculation Examination.—June 1869.

Friday, July 2.—Afternoon, 2 to 5.

1. Describe the experiment by which Lavoisier proved air to consist of oxygen and nitrogen. How many cubic centimètres of oxygen are contained in 100 cubic centimètres of air? How many cubic centimètres of hydrogen would be burnt by this quantity of oxygen?

2. Describe by an equation what takes place when metallic zinc is acted upon by diluted hydrogen sulphate (sulphuric acid). Taking the atomic weight of zinc as 65, what weight of hydrogen would be evolved, according to this process, by the solution of one kilogramme of zinc?

3. Describe an experiment by which you would prepare carbonic oxide gas from carbonic acid.

4. Give the names and chemical formulæ of the substances formed by the action of hydrogen nitrate (nitric acid) upon metallic copper. What happens when the gas produced in this reaction comes into contact with air?

5. How would you obtain evidence of the presence in a given solution of ammonia, whether free or combined?

6. How would you prepare sulphuretted hydrogen gas? What is formed when this gas is burnt in air?

7. What are the proportions by weight of the constituent elements of the following compounds:—Marsh gas; olefiant gas; carbonic acid; and water?

Answers to Matriculation Questions.

1. The proportion of oxygen by volume in 100 volumes of air is variously given. Some consider it to be 20.9, others 21. Taking it at 21, then as oxygen burns double its volume of hydrogen, 21 c. c. of oxygen will burn 42 c. c. of hydrogen.

2. $$\underset{62}{Zn} + H_2\,SO_4 = Zn\,SO_4 + \underset{2}{H_2}$$

65 grammes of zinc evolve from hydric sulphate 2 grammes of hydrogen

$$\therefore \frac{1}{32{\cdot}5} \text{ of a gramme evolves 1 gramme of hydrogen.}$$

(One kilogramme contains 1000 grammes.)

$$\therefore \frac{1000}{3{\cdot}25} \text{ grammes evolve 30.769 grammes of hydrogen.}$$

MATRICULATION EXAMINATION.

JANUARY, 1870.

Friday, January 14th.—Afternoon, 2 to 5.

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1. Describe by equations as many processes as you know for the preparation of oxygen gas.

2. What is meant by the term "hardness" as applied to water? How would you ascertain whether a particular sample of water is hard or not?

3. Explain by an equation the process of making ammonia by the action of slaked lime upon sal-ammoniac; and describe the properties of the gas.

4. A litre of carbonic acid is completely reduced to carbonic oxide by hot carbon. What volume of carbonic oxide is obtained? What volume of oxygen is needed for the complete combustion of this carbonic oxide?

5. A given sample of gas is found to have the property of kindling a glowing splint of wood. Enumerate the reactions by which you would decide whether it consists of oxygen or nitrous oxide.

6. A solid substance is suspected to contain some compound of carbon. How would you ascertain by experiment whether such is the case?

7. How would you prepare sulphurous acid gas? Describe and explain the reactions by which you would identify it.

MATRICULATION EXAMINATION.—JUNE, 1870.

FRIDAY, JULY 1.—AFTERNOON, 2 TO 5 O'CLOCK.

CHEMISTRY.

1. Express the temperature of 40° Fahrenheit on the centigrade scale; and express the temperature of —40° centigrade on the Fahrenheit scale.

2. Give the names and formulæ of the Oxides of Nitrogen, and calculate their percentage compositions.

3. The weight of 11.2 litres of Hydrogen, under normal conditions, being 1 gramme, what are the weights of 100 litres of Oxygen, Carbonic Oxide, Carbonic Acid, Chlorine, and Ammonia Gases respectively, under the same conditions?

4. What are the tests you would employ in order to detect the presence of a dissolved *Sulphate*, *Chloride*, *Sulphide*, and *Nitrate?*

5. What compounds can be formed by the combustion of the following substances, in a limited and in an excessive quantity of air; and what are their respective formulæ: — Sulphur, Phosphorus, Hydrogen, and Carbon?

6. Express in symbols what takes place (α) when Sulphuric Acid and Chloride of Sodium are heated together; (β) when Sulphuric Acid, Chloride of Sodium, and Binoxide of Manganese are heated together; (γ) when Nitrite of Ammonia and Nitrate of Ammonia are each of them subjected to heat.

7. How is Sulphuric Acid (Oil of Vitriol) made? Explain the chemical changes which take place in the process.

Matriculation Examination.

January, 1871.

Friday, January 13th.—Afternoon, 2 to 5.

1. Water is said to be a compound of hydrogen and oxygen. Describe experiments in proof of this view, and explain why hydrogen and oxygen are called elements.

2. Describe how you would prepare pure oxygen; and give an outline of the leading properties of the element.

3. What is the difference in composition between a sulphite and a sulphate, between a nitrite and a nitrate, and between nitrous oxide and nitric oxide?

4. A specimen of spring water is supposed to contain sulphuretted hydrogen (hydric sulphide). How would you ascertain whether any sulphuretted hydrogen was present?

5. Explain the action of oil of vitriol on common salt. What weight and what volume of hydrochloric acid gas (hydric chloride) could be obtained from 100 grammes of pure common salt (sodic chloride)? Explain the actions of aqueous hydrochloric acid on zinc, on red oxide of mercury (mercuric oxide), and on black oxide of manganese, respectively.

Answers to Questions and Calculations.

1. The answer to the first part of this question will be found in Chapter VII., and to the second part at page 2.

2. The answer to this question will be found in Chapter III.

3. The answer to the first part of this question will be found at pages 98 and 99, and to the second at pages 99 to 102.

4. The answer to this question will be found in the tests for hydric sulphide, at page 182.

5. The answer to the first part of this question will be found at pages 141 and 142.

$$Na = 23 \quad NaCl + H_2SO_4 = NaHSO_4 + HCl \quad H = 1$$
$$Cl = \frac{35.5}{58.5}\ (58.5) \qquad\qquad (36.5)\ Cl = \frac{35.5}{36.5}$$

∴ 58.5 parts of common salt give 36.5 of hydric chloride

1 part „ gives $\frac{36.5}{58.5}$ „

100 parts „ give $\frac{36.5 \times 100}{58.5}$ „

∴ 100 grammes „ give $\frac{36.5 \times 100}{58.5}$ grammes of hydric chloride

Ans. 63.39 grammes of hydric chloride.

38.5 grammes of hydric chloride occupy 22.4 litres

1 gramme „ occupies $\frac{22.4}{36.5}$ „

63.9 grammes „ occupy $\frac{22.4 \times 63.9}{36.5}$ litres

Ans. 38.29 litres.

The action of aqueous hydric chloride on zinc is to form zinc chloride and free hydrogen.

$$2\,(HCl) + M = MCl_2 + H_2$$

On mercuric oxide is to form mercuric chloride and water,

$$2\,(HCl) + HgO = HgCl_2 + H_2O$$

On manganic binoxide is to form manganous chloride, water, and perchloride,

$$4\,(HCl) + MnO_2 = MnCl_2 + 2\,(H_2O) + Cl_2$$

Matriculation Examination,

June, 1871.

Friday, June 30.—Afternoon, 2 to 5.

CHEMISTRY.

Examiners:—H. Debus, Esq., Ph.D., F.R.S.; and Prof. Odling, M.B., F.R.S.

1. When the rooms of a house are heated by means of hot water, or hot air, the heat is said to be transmitted by convection from the furnace to the rooms of the building. If you place your hand upon a piece of hot iron, the heat is communicated by conduction from the metal to your body. Explain what is meant by the terms "convection" and "conduction."

2. Why do you call oxygen, hydrogen, and carbon chemical elements? Describe the properties by which you would distinguish nitrous oxide (laughing gas) from a mechanical mixture of two volumes of nitrogen and one volume of oxygen?

3. What substance do you obtain when a piece of sulphur is burnt in oxygen? How can this substance be converted into common sulphuric acid? How would you ascertain whether a specimen of an acid consisted of sulphuric or of nitric acid?

4. Describe the structure of the flame of a common candle, and explain why carbon is set free by the combustion in the interior, but not on the surface, of the flame.

5. The hard water of our chalk districts is rendered

soft by addition of lime-water. Explain fully why lime-water is able to produce this result.

6. By what experiments could you prove that common phosphorus, and red or amorphous phosphorus, are allotropic modifications of the element phosphorus?

7. How much oxygen do you require for the complete combustion of 10 c.c. of marsh-gas? How many c.c. of each product of combustion would be formed?

Answer to question 7.

$$\underset{\text{2 vols.}}{CH_4} + \underset{\text{4 vols.}}{O_4} = \underset{\text{2 vols.}}{CO_2} + \underset{\text{4 vols.}}{2(H_2O)}$$

$\therefore$ 10 c.c. of marsh-gas require 20 c.c. of oxygen for their complete combustion. 10 c.c. of carbonic acid, and 20 c.c. of steam are formed.

MATRICULATION EXAMINATION.

JANUARY, 1872.

CHEMISTRY.

1. You find by experiment that the heat of a flame of a single spirit lamp is sufficient to boil 6 oz. of water at 100°C. in an open evaporating dish. Two flames having each the same size and temperature as the one just mentioned, are not able to render the water and escaping steam hotter than the single flame was able to render them. Why not?

How much steam at 100°C. is required to raise the temperature of 54 oz. of water from 0° to 100°C., the latent heat of steam being taken = 540?

2. An atmosphere is said to take part in the processes of Combustion, of Respiration, and of the growth and decay of Plants. Explain how, and by means of which of its constituents, the air acts on each of these processes.

3. Water is said to be composed of equivalent quantities of its two constituents, or of two atoms of hydrogen and one of oxygen; why do you consider two atoms of hydrogen to be equivalent to one of oxygen?

4. How would you prepare chlorine, and how do you explain your process? If one bottle contain chlorine gas and another bromine vapour, how could you dis-

tinguish the two substances from each other? Why can chlorine not exist free in nature?

5. Explain by means of symbols the formation of nitric acid from saltpetre and common sulphuric acid. You prepare hydrogen by the action of zinc upon dilute hydric sulphate; why cannot hydrogen be procured by the action of zinc upon dilute nitric acid? Explain what happens when concentrated nitric acid is placed in contact with a solution of potassic hydrate.

6. How much phosphoretted hydrogen, PH_3, by weight and by volume, would be furnished by 31 grammes of phosphorus, supposing all the phosphorus to be converted into the gas?

Answers to Matriculation Questions.

1. When water is boiled on the surface of the earth, its vapour, at the temperature of 100°C., is able to overcome the pressure of the atmosphere, and all the heat applied to the water at this temperature goes, not to raise the temperature of the water or steam, but to convert the water into steam. The quantity of heat applied to the water has no effect in raising its temperature above 100°C., if the pressure be that of about 15lbs. to the square inch, *i.e.* the atmospheric pressure.

Any weight of water converted into steam at 100°C. gives out, when the steam is condensed into water at 100°C., sufficient heat to raise the temperature of 450 times that weight of water one degree centigrade; therefore it can raise 45 times that weight of water through 100°C.—that is, one ounce of water converted into steam at 100°C., and condensed to water again at 100°C., can raise 45 ounces of water from 0° to 100°C.

2. The answers to this question will be found at pages 64 and 65, and pages 30 to 32. The decay of plants takes place in this manner. In the presence of moisture and a proper temperature the oxygen of the air unites with the carbon and other constituents of the plants, forming carbonic acid and water. When the plant is undergoing decomposition in this way, other products are formed by union of its constituents, such as marsh gas, but these are not required to be mentioned by this question.

3. Read at pages 209 and 210.

4. Read chapter on Chlorine, page 121, and at page 148. If this question requires you to distinguish between a chloride and a bromide (I cannot tell whether it does or not), you must add sodic or potassic hydrate to each substance, and shake them up till the chlorine and bromine are absorbed, then heat the products

in separate vessels to drivo off the oxygen, and you will have a bromide and a chloride left; then you can test as described at page 142, and page 150.

5. Pages 105 and 226—Page 109. It is not in this question mentioned whether the solution of potassic hydrate is strong or weak; if it were strong a very violent action would take place, but the products would bo the same. Hydric nitrate (Nitric acid, as it is called in the question) is an oxidising agent. By it the hydrogen would be oxidised into water, and some oxides of nitrogen would be given off, zinc sulphate being formed. This is why hydrogen cannot be evolved from zinc and dilute hydric nitrate.

6. $$P = 31, H_3 = 3, \therefore PH_3 = 34.$$

31 grammes of phosphorus unite with 3 grammes of hydrogen to form 34 grammes of phosphoretted hydrogen.

$$\therefore \text{ 34 grammes by weight are formed.}$$

$$PH = 2 \text{ vols.}$$

And 34 *grammes* of phosphoretted hydrogen measure 2 (11.2 litres.

Answers to Calculations.

No. 1.

Zero Fahrenheit is 32°F. below the zero Centigrade.

$\therefore$ 32 must be subtracted from 40°F.

$$40 - 32 = 8.$$

The number of Centigrade degrees between zero or freezing point and boiling point is 100. The number of Fahrenheit degrees between the freezing and boiling point is 180.

$$\therefore 100°C. = 180°F. \text{ or } 5°C. = 9°F.$$

$$\therefore 8°F. \text{ equal } \frac{8 \times 5}{9} = 4.\dot{4}C.$$

and 40°F. correspond to 4.4°C.

$$40°C. = \frac{40 \times 9}{5} = \frac{360}{5} = 72°F.$$

But Fahrenheit zero is 32° below freezing.

$\therefore$ from 72 degrees Fahrenheit below freezing the 32° above zero must be subtracted, $72 - 32 = 40$,

and so —40°C. correspond to —40°F.

The answer to the first part of this question will be found at page 223.

Answer to second part :—

Nitrous oxide.—Symbol $N_2O = 44$, $N = 28$, $O = 16$.

44 parts of nitrous oxide yield 28 of nitrogen.

$$\therefore 100 \text{ parts of it yield } \frac{28 \times 100}{44} = 63.6\dot{3} \text{ of nitrogen.}$$

100 parts — 63.6$\dot{3}$ of nitrogen leave 36.37 of oxygen.

Nitric oxide.—Symbol $NO = 30$, $N = 14$, $O = 16$.

30 parts of nitric oxide yield 14 of nitrogen.

$$\therefore 100 \text{ parts yield } \frac{14 \times 100}{30} = 46.\dot{6}.$$

100 parts — 46.6 leave 53.4 of oxygen.

Nitrous acid.—Symbol $N_2O_3 = 76$, $N = 28$, $O_3 = 48$.

76 parts of nitrous acid yield 28 of nitrogen.

$$\therefore \text{ 100 parts yield } \frac{28 \times 100}{76} = 36.\dot{8}.$$

100 parts — $36.\dot{8}$ leave 63.2 of oxygen.

Nitric peroxide.—Symbol $N_2O_4 = 92$, $N_2 = 28$, $O = 64$.

92 parts of nitric peroxide yield 28 of nitrogen.

$$\therefore \text{ 100 parts yield } \frac{28 \times 100}{92} = 30.44.$$

100 parts — 30.44 leave 69.56 of oxygen.

Nitric acid.—Symbol $N_2O_5 = 108$, $N_2 = 28$, $O_5 = 80$.

108 parts of nitric acid yield 28 parts of nitrogen.

$$\therefore \text{ 100 parts yield } \frac{28 \times 100}{108} = 25.92 \text{ of nitrogen.}$$

100 parts — 25.92 leave 74.08 of oxygen.

No 3.

11.2 litres of hydrogen weigh 1 gramme

1 litre ,, weighs $\frac{1}{11.2}$,,

100 litres ,, weigh $\frac{100}{11.2} = 8.928$ grammes.

Oxygen is 16 times as heavy as hydrogen.

$\therefore$ 100 litres weigh $8.928 \times 16 = 42.848$ grammes.

Carbonic oxide is 14 times as heavy as hydrogen.

The formula for the molecule of carbonic oxide is $CO = 28$. $CO = 28$ grammes occupies 2 (11.2) litres, and $\therefore$ 11.2 litres of carbonic oxide weigh 14 times as much as 11.2 litres of hydrogen.

$\therefore$ 100 litres weigh $8.928 \times 14 = 124.992$ grammes.

Carbonic acid is 22 times as heavy as hydrogen.

The formula for the molecule of carbonic acid is $CO_2 = 44$. $CO_2 = 44$ grammes occupies 2 (11.2) litres. $\therefore$ 11.2 litres of carbonic acid weigh 22 times as much as 11.2 litres of hydrogen.

$\therefore$ 100 litres weigh $8.928 \times 22 = 196.416$ grammes.

Chlorine is 35.5 times as heavy as hydrogen.

$\therefore$ 100 litres weigh $8.928 \times 35.5 = 316.944$ grammes.

Ammonia gas is 8.5 times as heavy as hydrogen, for the formula of molecule is NH_3 = 17. NH_3 = 17 grammes occupies 2 (11.2) litres. ∴ 11.2 litres of ammonia gas weigh 8.5 times as heavy as 11.2 litres of hydrogen.

∴ 100 litres weigh 8.928 × 8.5 = 75.888 grammes.

4. Answers to be found for a *Sulphate* at page 179. For a *Chloride* at page 142. For a *Sulphide* at page 182. For a *Nitrate* at page 110.

5. *Sulphur* at pages 170 and 242. *Phosphorus* at pages 188, 193, 249, and 252. *Hydrogen* at pages 39 and 218. *Carbon* at pages 71, 219, and 220.

6. (α) At pages 138 and 237; (β) at pages 123 and 233; (γ) at pages 218 and 224.

7. *Sulphuric acid*—under the head of *Hydric sulphate* (its scientific name)—at pages 175 and 244.

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